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A SHORT COURSE

IN

QUALITATIVE ANALYSIS

WITH THE NEW NOTATION.

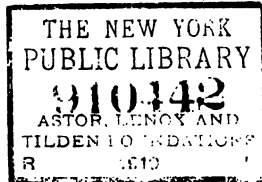
BY
J. M. CRAFTS,

PROFESSOR OF GENERAL CHEMISTRY IN THE CORNELL UNIVERSITY.

SECOND EDITION.

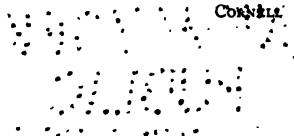
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This Book is inscribed by the Author

TO

M. ADOLPH WURTZ,

AS A TOKEN OF AFFECTIONATE REGARD, TO A FRIEND AND TEACHER,

AND AS A TRIBUTE OF RESPECT AND ADMIRATION, TO THE

MASTER OF A SCHOOL IN MODERN CHEMISTRY.

PREFACE.

THIS little work was written for the use of a class of students in the Cornell University, who take a year's course of chemistry, including four hours a week of laboratory practice; reference was also had to the requirements of the scientific students in Union College, whose course is nearly equivalent to that mentioned. The author is indebted to the kindness of Professor Perkins for many valuable suggestions, and for the compilation of Tables IV and V at the end of the book.

A considerable portion of the introductory part of this book is devoted to an explanation of the theory of chemical reactions and nomenclature. Many of the standard works on analytical chemistry still use the old notation, and the formulas to be found in them do not correspond to those used in the best text-books on general chemistry; and none of the elementary works for the laboratory-student supply the want, often felt by him, of a system of rules, at hand for use at the moment when he most requires them, namely, when he is writing the formula of a reaction at his desk in the laboratory with his tests before him. The author has had these points in view in writing the two introductory chapters.

It might be objected that the theory of chemical notation should be found in text-books on general chemistry ; but even when the student has mastered the rudiments of the science, as given in any of the best modern works, he will find the arrangement of many of them inconvenient for reference, although it is excellent for instruction, and, moreover, it is by no means necessary that the study of those works should precede that of analytical chemistry.

It is quite feasible to commence a course of experiments in the laboratory concurrently with the study of any of the elementary text-books on general chemistry, or with the attendance upon lectures, illustrated with the usual experiments. Chemistry certainly becomes a more attractive study, when the practical and the theoretical present themselves side by side, so that while the theory explains the experiment, the experiment awakens an interest in the theory ; and no course of study is more apt to interest the beginner in chemistry than that of the admirably simple and delicate tests of qualitative analysis ; tests which illustrate the general laws of the science, while they have a very direct bearing upon some of the problems of everyday life.

Analytical chemistry, besides its immediate value as an important branch of knowledge, cannot be too highly prized as affording a convenient introduction to the methods of investigation used in an experimental science, and as offering a means of education of many faculties, which are not easily developed by school or university training. The importance of laboratory experiments is awakening every day increased attention, and the time is fast passing by, when chemistry is taught to persons, who suppose that they have a vocation for a scientific profession, only by lectures and recitations.

The system of analysis, given in Part III, is founded upon that of Fresenius, and includes a minute description of all the steps to be taken in performing tests and directions for passing from one test to another; and since these details, which are required by the beginner, not only become unnecessary, after a certain familiarity with analytical work has been acquired, but also, under the scheme less convenient for reference, tables similar to those of Will have been added, which indicate the important tests and leave the experience of the student to suggest the proper mode of applying them.

The Tables IV and V are intended to record a number of facts in analytical chemistry in a compact form, and to give an exact conception of what is meant by the insolubility of a precipitate, and a means of judging of the advantages of different methods of precipitation.

ITHACA, *August* 19, 1869.

PREFACE TO THE SECOND EDITION.

A few changes have been made in the second edition, which have been suggested by the use of the previous one in the laboratory and some practical directions have been added, which are intended to mark out more closely the course of study, which it seems most advisable to follow.

One of the principal applications of analytical chemistry is to the investigation of minerals, and it may not be out of place here to recommend to the consideration of instructors and students the combination of mineralogical work with laboratory practice. A convenient order of study is, first to acquire familiarity with the methods of detection and separation of bodies in solution, which are given in Part III, and then, before taking up the analysis of solid bodies, to study their crystallographical and general mineralogical characters from any elementary work treating of the subject. Turning again to the general analysis of solid bodies (minerals and technical products of all kinds), the course should be made as varied and extended as possible; and it will be found that the drill in blowpipe practice

and the study of the physical characters of minerals make an excellent preparation for the preliminary testing of solids, Part III, and often afford hints, which are of great value in the subsequent performance of the complete analysis in the wet way.

It is quite possible for a class to complete a course like the one proposed by working four hours a week for a year in the laboratory.

ITHACA, 1870.

A SHORT COURSE IN ANALYTICAL CHEMISTRY.

PART I—INTRODUCTION.

CHAPTER I.

The analytical chemist is not required to investigate a large range of substances, and the theory of the composition of those bodies, which he usually deals with, can be briefly explained.

ELEMENTS. The ultimate result of the analysis or decomposition of all matter is the discovery of a number of substances which cannot be decomposed, and are therefore called simple or elementary substances. The table on the next page contains the names of 32 elements. They are chosen from among the 64 elements which have been discovered, because they are more frequently met with than the others, and because they are the only ones treated of in this work.

The task of the analytical chemist is to recognize these elements, when they occur alone, or to separate them from the compounds of which they form a part, or most frequently to decompose a compound substance into simpler compounds, and to isolate the latter in such a form that they can be easily recognized. When the different elements are known, which

NON-METALLIC ELEMENTS.					
MONADS.	DYADS.	TRIADS.	TETRADS.	PENTADS.	HEXADS.
Fluorine, F... 19. Chlorine, Cl.. 35.5	Oxygen, O 16. Sulphur, S..... 32.	Nitrogen, N.. 14. Boron, B..... 11.		Nitrogen, N.. 14. Phosphorus, P 31.	Sulphur, S..... 32.
Hydrogen, H. 1.			Silicon, Si..... 28. Carbon, C..... 12.		
		Arsenic, As.... 75. Antimony, Sb 122. Bismuth, Bi ...210.	Platinum, Pt.... 197.4 Tin, Sn 118.	Arsenic, As... 75. Antimony, Sb. 122. Bismuth, Bi ...210.	
Silver, Ag.....108.	Lead, Pb.....207. Copper, Cu..... 63.4 Mercury, Hg.. 200.	Gold, Au197.			
	Nickel, Ni..... 58.8 Cobalt, Co..... 58.8 Iron, Fe..... 56. Manganese, Mn. 55. Zinc, Zn..... 65.2		Iron, Fe 56. Manganese, Mn. 55. Chromium, Cr. 52.5 Aluminium, Al. 27.4		Manganese, Mn. 55. Chromium, Cr.. 52.5
Sodium, Na... 23. Potassium, K. 39.1	Magnesium, Mg. 24. Calcium, Ca 40. Barium, Ba137.				

METALS.

NON-METALLIC
ELEMENTS.

constitute such simpler compounds, the composition of the complex body from which they were obtained can be determined.

All substances are formed by the union of chemical elements, and a previous study of the manner in which the elements combine with each other is essential to the successful pursuance of analytical investigations.

The language in which chemists express their ideas regarding the chemical constitution of bodies comprises certain symbols or abbreviations, to which a conventional meaning is attached, and formulas, which are made by grouping symbols together.

SYMBOLS. The symbols which stand in the table after the names of the elements are abbreviations, which are used instead of the names in writing chemical formulas.

CHEMICAL FORMULAS describe by means of symbols the chemical constitution of bodies.

Example:— HCl is the formula of hydrochloric acid, and signifies that it is composed of hydrogen and chlorine.

THE COMBINING WEIGHTS are the numbers standing in the table after the symbols. They are peculiar to each element, and denote the proportion by weight in which it unites with other elements. In formulas the symbols stand for these weights, as well as for the names of the elements; thus, in the formula of hydrochloric acid, HCl signifies that 1 part by weight of hydrogen is combined with 35.5 parts by weight of chlorine.

ELEMENTS COMBINE WITH EACH OTHER IN NO OTHER PROPORTIONS BY WEIGHT THAN THOSE EXPRESSED BY THE COMBINING NUMBERS (or by very simple multiples of them). This statement is called the **LAW OF DEFINITE PROPORTIONS**.

USUALLY THE MULTIPLES OF THE COMBINING NUMBERS: 2, 3, 4, 5, 6, 7, EXPRESS THE PROPORTIONS IN WHICH THE ELEMENTS COMBINE WITH EACH OTHER. This statement is called the **LAW OF MULTIPLE PROPORTIONS**.

Example:—Nitrogen combines with oxygen only in the proportions expressed by the formulas: N_2O , NO , N_2O_3 , NO_2 , and N_2O_5 , or, referring to the table for the combining numbers

which the symbols stand for, twice 14 parts of N unite with 16 parts of O; 14 parts of N with 16 parts of O; twice 14 parts of N with 3 times 16 parts of O; &c.

CHEMICAL ATOMS. It is supposed that the utmost limit to which the division of matter could be carried, would lead to its separation into a great number of particles, so small as to be incapable of further division. With reference to their quality of indivisibility, such particles of matter are called atoms (from *α*, privative, and *τέμνω*, I cut). Atoms, therefore, are the indivisible constituents of matter. It is further supposed, that chemical combination consists in the union of atoms, or groups of atoms, and chemical decomposition in the separation of atoms or groups of atoms; and a chemical change supposes a change in the arrangement or grouping of the atoms of a body, involving the destruction of the previous arrangement.

ATOMIC WEIGHTS. The atomic theory attaches a new meaning to the combining weights of the elements, and defines them as the relative weights of atoms. Thus, if the weight of an atom of hydrogen is $1 \times w$ ($w =$ a very small unknown quantity), the weight of an atom of chlorine is $35.5 \times w$. Not the absolute weights of atoms but their relative weights, have been discovered. The weight of an atom of hydrogen is taken as a standard, and called 1; hence the weight of the atoms of other elements are expressed in terms of this unit. Example, 35.5 for chlorine, 16 for oxygen. With reference to the above theory, the combining weights are usually called atomic weights. The full meaning, therefore, of the chemical formula HCl is, that the body which it represents consists of compound atoms, each one containing an atom of hydrogen and an atom of chlorine. An atom of chlorine weighs 35.5 times as much as an atom of hydrogen, so that the proportion by weight of each constituent of the body is expressed by its formula.

CHEMICAL NOTATION. Any change in the constitution of bodies, as well as their formation and decomposition, involves what is called a chemical reaction.

A CHEMICAL REACTION may be described as a change in the arrangement, or the state of combination of the atoms of bodies. Such a change can be denoted by combining formulas together in the same way that quantities are combined in common algebraic calculations. The signs used are $+$, $-$ and $=$. Coefficients are only used to multiply the symbols to which they are joined. When placed on the line, they multiply all the symbols which follow. When placed below or above the line, they are used to multiply only the symbol, or the group of symbols in brackets, that immediately precedes them. Brackets are used to distinguish certain groups of atoms in a compound from the remaining atoms. The combination of atoms is expressed by writing their symbols side by side, or by grouping them together without $+$ or $-$. The sign $+$ expresses that the bodies connected by it are brought in contact with each other by addition, but that they are not combined.

Example:— $\text{Ba}(\text{NO}_3)_2 + \text{Ca SO}_4$. The formula indicates that a compound $\text{Ba}(\text{NO}_3)_2$ containing barium, nitrogen, and oxygen, in the proportion of 1 atom of barium combined with twice 1 atom of nitrogen and twice 3 atoms of oxygen, is brought in contact with a compound $[\text{Ca SO}_4]$ containing calcium, sulphur, and oxygen, in the proportion of 1 atom of calcium combined with 1 atom of sulphur and 4 atoms of oxygen. The result of a reaction is denoted by combining the formulas of the bodies, which take part in it, as in an ordinary equation. The formulas before the sign ($=$) indicate the state of combination of the atoms before the reaction; those after the sign of equality ($=$) show the state of combination of the atoms after the reaction.

Example:— $\text{Ba}(\text{NO}_3)_2 + \text{Ca SO}_4 = \text{Ba SO}_4 + \text{Ca}(\text{NO}_3)_2$.

The equation expresses the result of bringing in contact the bodies described above, viz.: the formation of new compounds, containing barium, sulphur and oxygen, and calcium, nitrogen and oxygen.

CHEMICAL AFFINITY.

The force which impels atoms to unite with other atoms is called chemical affinity. The quantity or the nature of the force inherent in the atoms of every substance determines the chemical properties of the substance. The study of the results of the action of chemical affinity is the province of chemistry.

The phenomena, which the action of chemical affinity gives rise to, can best be studied under several heads.

Firstly.—Chemical affinity may cause the atoms of an elementary body to unite among themselves. Only the cases of such action in which the element is capable of assuming the simplest physical condition of matter, namely, the form of a gas, have been studied satisfactorily.

The following conclusions in regard to the state of combination of the atoms of elementary bodies have been arrived at:—

The atoms of mercury and zinc remain separate in the gaseous state.

The atoms of hydrogen, oxygen, chlorine, bromine, iodine, nitrogen (and sulphur at a temperature higher than 1000° centigrade) are united in groups of two atoms.

The atoms of phosphorus and arsenic in the gaseous state unite in groups of four atoms. Sulphur, at a temperature of 500° centigrade, in groups of six atoms.

The formulas for these bodies in a gaseous state are:

Hg	HH or H_2	PPPP or P_4	SSSSSS or S_8
Zn	OO or O_2	AsAsAsAs or As_4	
	ClCl or Cl_2		
	BrBr or Br_2		
	NN or N_2		
	SS or S_2		

MOLECULE. At this point a definition of the term molecule is required. A molecule is the smallest particle of a body which can exist alone.

The molecules of elementary bodies contain one or more at-

oms of the same kind. The molecules of compound bodies contain two or more atoms of different kinds.

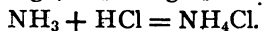
A molecule of mercury, hydrogen, hydrochloric acid, or water is represented by the respective formulas, Hg, H₂, HCl, or H₂O.

ATOM. An atom may be further defined as the smallest particle of matter which can take part in a chemical reaction. Atoms, therefore, appear while a chemical reaction is going on, although it is impossible to suppose that a physical subdivision of matter could be carried further than the isolation of molecules.

Thus in the reaction: $\text{Na}_2\text{S} + \text{CuCl}_2 = \text{CuS} + 2\text{NaCl}$, the force of chemical affinity breaks up the molecules, (Na₂S) and (CuCl₂) to form the new ones, (CuS) and (NaCl), and during this reaction the atoms Na, S, Cu and Cl, must be set free from their combinations, and therefore must exist as atoms.

Secondly.—Chemical affinity may combine atoms of a single element, or groups containing atoms of one or more elements, with the atoms of another element, or with groups containing atoms of one or more other elements.

Examples:— $\text{Hg} + \text{Cl}_2 = \text{HgCl}_2$.

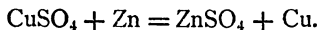


The inverse action frequently takes place through the agency of heat or of some other force, and groups of atoms (molecules) break up into other groups, which are usually simpler in constitution than the primitive ones.

Example:—Hg (CN)₂ when heated becomes Hg + (CN)₂.

Thirdly.—Chemical affinity may cause compound bodies, brought in contact with each other, to mutually exchange some of their constituents; or an atom or a group of atoms may substitute itself for another atom or for a group of atoms in a compound body.

Examples:— $\text{Na}_2\text{CO}_3 + \text{BaCl}_2 = \text{BaCO}_3 + 2\text{NaCl}$.



THE QUALITY OF THE CHEMICAL AFFINITY [inherent in the atoms of each element determines the part which the element

will play in the different chemical changes mentioned above. It is usually necessary to study each particular case, in order to determine the exact result of bringing in contact any two substances. Empirical rules, however, defining the nature of the chemical affinity of the elements and the consequences of its action, can be given in a few cases. These rules are not capable of a very strict application, but they serve to indicate in most cases, when a number of bodies are brought together in a reaction, those which will probably combine with each other. Gold is attacked by acids less readily than the metals of the arsenic group. The difference between the affinities of the metals which are ranged in the same group in the table (page 2) is too slight to be of consequence in the application of rule second.

First.—COMBINATION USUALLY OCCURS BETWEEN METALS AND NON-METALLIC ELEMENTS, LESS READILY BETWEEN DIFFERENT NON-METALLIC ELEMENTS, AND LEAST READILY BETWEEN DIFFERENT METALS. See table of the elements, page 2.

The following non-metallic elements or oxygen compounds of non-metallic elements, combined with hydrogen, form acids, and combined with metals form salts. STRONG ACIDS—sulphuric, H_2SO_4 ; nitric, HNO_3 ; chloric, HClO_3 ; chlorhydric, HCl . WEAK ACIDS—sulphurous, H_2SO_3 ; chromic, H_2CrO_4 ; phosphoric, H_3PO_4 ; boracic, H_3BO_3 ; oxalic, $\text{H}_2\text{C}_2\text{O}_4$; acetic, $\text{HC}_2\text{H}_3\text{O}_2$; fluorhydric, HF ; sulphydric, H_2S ; cyanhydric, HCN ; carbonic, H_2CO_3 ; silicic, H_4SiO_4 . A metal has a tendency to substitute itself for the hydrogen in a strong acid, to form a salt with it, in preference to a weak one; so that a strong acid usually displaces a weak one from its salts.

Second.—THE METALS WHICH STAND LOWEST IN THE TABLE (page 2) HAVE THE GREATEST TENDENCY TO COMBINE WITH THE STRONGEST ACIDS.

Example:— $\text{CuSO}_4 + \text{Zn} = \text{ZnSO}_4 + \text{Cu}$.

Third.—Another rule which sometimes takes precedence of the second, is, that WHEN FROM SOME OF THE CONSTITUENTS OF DIFFERENT COMPOUNDS IN SOLUTION, AN INSOLUBLE BODY CAN

BE FORMED, THE ELEMENTS WHICH WOULD COMPOSE SUCH A BODY GENERALLY UNITE WITH EACH OTHER.

Example:— $\text{Na}_2\text{CO}_3 + \text{Ca}(\text{HO})_2 = 2\text{NaHO} + \text{CaCO}_3$. The compound CaCO_3 is formed in virtue of its insolubility, when the reaction takes place in an aqueous solution.

ATOMICITY.

The quality of the chemical affinity of the elements determines the nature of reactions. The quantity of their chemical affinity determines the proportions in which the elements combine with each other. The result of the combination of two or more atoms with each other is the neutralization of the chemical force which brings them together, so that usually no force is left in the compound tending to combine other bodies with it. Thus a molecule of HCl has no power to combine further with atoms of H or of Cl . The atoms of some elements are animated with greater quantities of chemical force than those of other elements. Thus an atom of oxygen may unite with an atom of hydrogen, and still be capable of combining with another atom of hydrogen, or with an atom of chlorine. If an atom of hydrogen has one unit of chemical force, an atom of oxygen has two units, carbon has four, and nitrogen has five. The number of units of chemical force residing in an atom is called its ATOMICITY. The table of the elements, page 2, classifies them according to their atomicity, into monoatomic elements, or monads; and diatomic elements, or dyads, &c.

Some of the units of chemical force of an element may lie dormant until developed by the approach of some other force, which awakens them. Iron, for instance, acts as a dyad when no element is present to call forth all of its four units of chemical force. It is worthy of note, that in such cases two units of force generally disappear together, as if they became dormant by each one neutralizing the effect of the other. It will be noticed that in the table, page 2, sulphur is placed in the column of dyads, and also in that of the hexads; this is because in many cases four of the six units of chemical force in sulphur lie

dormant, and the element plays the part of a dyad. In nickel, cobalt, and iron, two of the four units of chemical force frequently lie dormant, and hence these metals often act as dyads. Only the atomicity of each element, which is usually displayed in the kind of reactions which we have to consider, is cited in the table.

The knowledge of these facts is essential to enable the student to write formulas correctly.

Examples:—The compound of barium and chlorine must contain two atoms of chlorine, combined with one atom of barium. Its formula is BaCl_2 . The greatest amount of oxygen, that an atom of carbon can unite with, is two atoms. The formula of the compound is CO_2 . When zinc is substituted for silver in a compound, one atom of zinc, with its two units of chemical force or affinity, takes the place of two atoms of silver, because each atom of silver has only one unit of chemical force. Thus, in writing formulas, an atom of one element is equivalent to or takes the place of another element of the same class. In comparing elements of different classes, their value in an equation depends upon the number of units of chemical force which they contain.

TWO MONADS ARE EQUIVALENT TO A DYAD; THREE MONADS TO A TRIAD, ETC. THREE DYADS ARE EQUIVALENT TO TWO TRIADS, ETC.

Examples:— $\text{PbO} + 2\text{HCl} = \text{PbCl}_2 + \text{H}_2\text{O}$.

$\text{SiCl}_4 + 2\text{H}_2\text{O} = \text{SiO}_2 + 4\text{HCl}$.

ATOMICITY OF GROUPS OF ATOMS.—When part of the affinities or units of chemical force of a polyatomic atom are satisfied or neutralized, the residual atomicities, or those which remain free, determine the atomicity of the group.

Example:—Nitrogen, a pentad, when combined with three atoms of hydrogen, is capable of uniting with two other monads or with one dyad— NH_3 can unite with H and Cl to make NH_4Cl . Moreover, certain groups of elements take part in many chemical reactions without being broken up, and, so far as this.

is the case, they may be considered as playing the part of elementary bodies in the reactions. Such groups are frequently called radicals, with reference to a theory that they are the roots of compounds.

The following table, showing the atomicities of such groups, will be found convenient in writing formulas :

Monads.	Dyads.	Triads.	Tetrads.	Hexads.
HO*	SO ₄ †	PO ₄	SiO ₄	Fe ₂ §
NO ₃ †	SO ₃	BO ₃	Fe(CN) ₆	Cr ₂
ClO ₃	CrO ₄			Al ₃
C ₂ H ₃ O ₂	CO ₃			Fe ₂ (CN) ₁₂
NH ₄	C ₂ O ₄			
CN	C ₄ H ₄ O ₆			
	Hg ₂			

Examples :—Na(HO); H(NO₃); (NH₄)(HO); Pb(C₂H₃O₂)₂;

* The group HO is a monad because one of the two affinities of the oxygen atom is satisfied and nullified by the affinity of the hydrogen with which it is combined, leaving one affinity free in the oxygen atom.

† The existence of the group NO₃, might seem inconsistent with the five-atomic character of nitrogen. The explanation is, that four affinities in the nitrogen atom are used to unite to it two atoms of oxygen, and the fifth combines with only one of the affinities of a third atom of oxygen, leaving one oxygen affinity free, which determines the mono-atomic character of the group.

‡ A similar argument applies to the case of the group SO₄. Here, sulphur being hexatomic, four affinities (atomicities) are used to combine with those of two atoms of oxygen, while the two remaining affinities of the sulphur atom are each combined with one in each of the two remaining atoms of oxygen, leaving two oxygen-affinities free.

§ In the case of Fe₂, &c., two of the eight affinities, belonging to two atoms of iron, are used to bind the two atoms together, leaving six free, and for this reason the group Fe₂ is six-, and not eight-atomic. It may be noticed, that groups of this nature, having free chemical affinities, can only be found in the case of polyatomic elements, because, when two mono-atomic elements combine with each other, no chemical affinity can be left free. For fuller explanations of the laws of chemical combination, see Frankland's Lecture Notes for Chemical Students.

$\text{H}_2(\text{SO}_4)$; $\text{Ba}(\text{CrO}_4)$; $\text{Hg}_2(\text{Cl}_2)$; $\text{H}_3(\text{PO}_4)$; Fe_2Cl_6 ; Cr_2O_3 ;
 $\text{Al}_2(\text{HO})_6$.

IN WRITING FORMULAS, THE SYMBOL OF THE ELEMENT, OR GROUP OF ELEMENTS, HAVING MOST DECIDEDLY THE CHARACTER OF A METAL, IS PLACED FIRST.

Example :— $\text{NaNH}_4\text{HPO}_4$.

CHAPTER II.

CHEMICAL NOMENCLATURE.

THE names of chemical compounds, which are used in works on analytical chemistry, do not describe their constitution as fully as formulas do; they serve, however, to identify the bodies and to recall certain principles of classification.

In the ordinary operations of analysis, the only bodies which are dealt with are binary compounds (i. e. compounds which only contain two elements, as KCl and H_2O) and compounds which contain a non-metallic element and oxygen combined with a metal or with hydrogen, as BaSO_4 , H_3PO_4 .

THE FORMATION OF THE NAMES OF COMPOUNDS.* BINARY COMPOUNDS FORM THEIR CLASS-NAMES OR GENERAL NAMES BY JOINING THE TERMINATION *ide* TO THE NAME OF THE PRINCIPAL NON-METALLIC ELEMENT WHICH THEY CONTAIN. *Examples*.—Chlorides, from chlorine; sulphides, from sulphur.

COMPOUNDS CONTAINING A NON-METALLIC ELEMENT AND OXYGEN COMBINED WITH A METAL FORM THEIR CLASS-NAMES BY JOINING THE TERMINATION *ate* OR *ite* TO THE NAME OF THE NON-METALLIC ELEMENT. *Ate* IS USED FOR THE PRINCIPAL COM-

* These rules refer specially to compounds of non-metallic elements with metals. When a compound consists of two non-metallic elements usage differs in regard to the one which shall give the class-name to the compound. The only cases of this kind, which are of importance, are treated separately under the head of oxides and acids, page 14 and page 15.

POUND WHICH CONTAINS THE MOST OXYGEN; *ite* FOR THE PRINCIPAL COMPOUND WHICH CONTAINS THE LEAST OXYGEN.

Examples :—Chlorates take their class-name from the group, ClO_3 ; chlorites from the group, ClO_2 .

THE PREFIXES *per* AND *hyper* ARE USED TO DENOTE HIGHER COMBINATIONS AND *hypo* TO DENOTE LOWER COMBINATIONS.

Examples :—Hyperchlorates take their name from the group, ClO_4 (chlorates from ClO_3); and hypochlorites from ClO (chlorites from ClO_2).

* THE METALLIC ELEMENT (OR WHEN THE COMPOUND ONLY CONTAINS NON-METALLIC ELEMENTS, THE ONE WHICH IS CONCEIVED TO GIVE IT ITS SPECIFIC CHARACTER) GIVES THE SPECIFIC NAME TO A COMPOUND AND TAKES THE ADJECTIVE TERMINATION *ic* OR *ous*, OR IT IS USED AFTER THE PREPOSITION *of*. *ic* DENOTES THE HIGHER STAGE OF COMBINATION; *ous* THE LOWER.

The termination *ic* is used for the most important combination which contains a large proportion of the element giving the class-name to the compound; the termination *ous* for that which contains a smaller proportion of the element giving the class-name. The prefixes *per*; *hyper* and *hypo* are used for still higher and still lower stages of combination.

Examples :—Oxide of zinc, ZnO ; sodic chloride or chloride of sodium, NaCl ; stannic chloride, SnCl_4 ; stannous chloride, SnCl_2 ; ferric sulphate, $\text{Fe}_2(\text{SO}_4)_3$; ferrous sulphate, FeSO_4 ; nitrous oxide, N_2O ; nitric oxide, NO ; hypernitric oxide, NO_2

WHEN ONLY ONE STAGE OF COMBINATION EXISTS, THE TERMINATION *ic* IS ALWAYS USED. Prefixes with a numerical signification are often used to distinguish between different states of combination in the place of the change of the adjective termination.

Example :—Oxide of lead, PbO ; binoxide of lead, PbO_2 .

OXIDES. The preceding rules suffice for the nomenclature of the oxides of metals, but a certain class of the oxides of non-metallic elements are frequently named according to a different system.

If an oxide is capable of combining with water to form an acid, it is called an anhydride or anhydrous acid (acid deprived of water), or sometimes, by way of abbreviation, an acid.

Examples.—Sulphuric anhydride, SO_3 —anhydrous sulphurous acid, or sulphurous acid, or sulphurous anhydride, SO_2 .

Many of the compounds of the non-metallic elements with each other are only known by irregular or common names as:—

Hydric oxide, H_2O = water.

Hydric nitride, NH_3 = ammonia.

Hydric arsenide, AsH_3 = arseniuretted hydrogen.

ACIDS. A compound of a non-metallic element with hydrogen is called an acid, when it possesses the property of readily exchanging its hydrogen for a metal, to form a salt. The most important of this class of acids are chlorhydric acid, HCl ; sulphydric acid, H_2S ; and fluorhydric acid; HF .

Example.— $2\text{HCl} + \text{Zn} = \text{ZnCl}_2 + 2\text{H}$.

The names of acids of this class always contain the radical *hydr* and the termination *ic*.

OXYGEN ACIDS. Another very important class of acids is that which contains a compound of hydrogen with a non-metallic element and oxygen. The names of such acids are derived from the characteristic non-metallic element, with the termination *ic* or *ous*, and, if necessary, the prefixes *hypo* and *hyper*, or *per*. The presence of oxygen is only denoted by the terminations or prefixes used in the same sense as on page 13.

Examples.— H_2SO_4 = sulphuric acid.

H_2SO_3 = sulphurous acid.

$\text{H}_2\text{S}_2\text{O}_3$ = hyposulphurous acid.*

SALTS. When in the place of hydrogen a metal is combined with a non-metallic element, or with a non-metallic

* Here the prefix *hypo* denotes, not that the substance contains less oxygen, but that the quantity of oxygen relatively to that of the sulphur is less than in sulphurous acid.

element and oxygen, the compound is called a salt. Common salt, NaCl, the chloride of sodium, is considered the type of this class of bodies.

The terminations of the names of salts follow the rules, pages 13 and 14. When a salt contains but one non-metallic element its termination is in *ide*. *Example*: K_2S , potassic sulphide. When a salt is derived from an oxygen acid, the termination of the acid corresponding to it is changed from *ic* to *ate*, and from *ous* to *ite*.

Examples:— H_2SO_4 , sulphuric acid; Na_2SO_4 , sodic sulphate; H_2SO_3 , sulphurous acid; Na_2SO_3 , sodic sulphite. The name of the metal follows the rules already given. (See page 14.)

Examples:— $Hg_2(NO_3)_2$, mercurous nitrate; $Hg(NO_3)_2$, mercuric nitrate.

It will be seen by the above examples that similar terminations designate different relative proportions of the elements in different cases. Thus, for sulphur, salts containing SO_4 take the termination *ate*. For nitrogen, salts containing NO_3 take the same termination, because in each case these are respectively the highest important oxygen compounds of each element. The meaning, therefore, of these terminations for each element must be committed to memory.

HYDRATES. The monoatomic group of elements HO combines with metals to form a class of bodies called hydrates.

Examples:— $Na(HO)$, sodic hydrate.

$Fe(HO)_2$, ferrous hydrate.

$Al_2(HO)_6$, aluminic hydrate.

It will be seen by the foregoing examples how much more fully the formulas express the composition of bodies than the names.

The following list of chemical compounds contains the names and formulas of the substances most frequently used for testing in the laboratory. These names and formulas should be committed to memory, and the rules of nomenclature may be studied in their application to them:

Chlorhydric, or hydrochloric ↓ acid, HCl ;	Potassic ferrocyanide, $\text{K}_4(\text{FeCy}_6)$, or $\text{K}_4\text{Fe}(\text{CN})_6$;
↓ Nitric acid, HNO_3 ;	Potassic ferricyanide,
↓ Sulphuric acid, H_2SO_4 ;	$\text{K}_6(\text{Fe}_2\text{Cy}_{12})$, or $\text{K}_6\text{Fe}_2(\text{CN})_{12}$;
↓ Sulphydric acid, H_2S ;	Potassic sulphocyanate,
Acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$;	$\text{K}(\text{CyS})$, or KCNS ;
Ammonic hydrate, NH_4HO ;	Calcic hydrate, $\text{Ca}(\text{HO})_2$;
Ammonic sulphide, $(\text{NH}_4)_2\text{S}$;	Calcic chloride, CaCl_2 ;
Ammonic carbonate, $(\text{NH}_4)_2\text{CO}_3$;	Calcic sulphate, CaSO_4 ;
Ammonic chloride, NH_4Cl ;	Baric chloride, BaCl_2 ;
Ammonic Oxalate, $(\text{NH}_4)_2\text{C}_2\text{O}_4$;	Baric nitrate, $\text{Ba}(\text{NO}_3)_2$;
Ammonic molybdate, $(\text{NH}_4)_2\text{MoO}_4$;	Baric carbonate, BaCO_3 ;
Sodic hydrate, NaHO ;	Magnesian sulphate, MgSO_4 ;
Sodic carbonate, Na_2CO_3 ;	Ferrous sulphate, FeSO_4 ;
Disodic hydric phosphate, Na_2HPO_4 ;	Ferric chloride, Fe_2Cl_6 ;
Sodic acetate, $\text{NaC}_2\text{H}_3\text{O}_2$;	Cobaltic nitrate, $\text{Co}(\text{NO}_3)_2$;
Potassic dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$;	Plumbic acetate; $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$;
	Argentive nitrate, AgNO_3 ;
	Mercuric chloride, HgCl_2 ;
	Platinic chloride, PtCl_4 ;
	Stannous chloride, SnCl_2 ;
	Alcohol, $\text{C}_2\text{H}_6\text{O}$.

After having made himself familiar with the formulas of the substances with which he tests, the student should write in the form of an equation the result of the action of each test, which he performs upon a compound under examination.

CHAPTER III.

GLASS APPARATUS.

The most important pieces of apparatus required by a student for the methods of analysis which follow, are mentioned in alphabetical order.

1. BOTTLES. 3—4 wide-mouthed bottles for corks, of 2—4 and 8 oz. capacity.

2. BLUE GLASS. A piece of glass 3 in. square, colored blue by cobalt.

3. FLASKS. Several 4 and 8 oz. Bohemian glass, flat-bottomed flasks. Beaker glasses may be used instead of flasks, but they are more expensive.

4. FUNNELS. 3—4 glass funnels from 1 inch to 3 inches diameter. Those with an angle of 60° should be chosen.

5. RODS. 2 glass rods 6 inches long. The glass should be cut of the proper length with a file, and the edges should be rounded by heating them with the blowpipe.

6. TEST-TUBES can be purchased of any dealer in chemical apparatus. 7 tubes 4 inches long and 6 tubes 5 inches long make a convenient set. These tubes are so thin and well annealed that a liquid may be heated in them directly over the flame of the lamp, without fear of breakage, provided the tube is moved rapidly up and down in the flame.

7. TEST-TUBE RACK. A rack of any form to hold 13 test-tubes is necessary. It is advisable to provide it with wooden pins, upon which the tubes may be inverted and left to dry.

8. TUBES. $\frac{1}{4}$ lb. rather thin German glass tubing, of two sizes: $\frac{3}{8}$ inch diameter and $\frac{1}{8}$ inch diameter.

9. WASH-BOTTLE. The wash-bottle is used to force a fine jet of water through a glass tube in order to direct it upon any object which it is desired to wash. A piece of $\frac{1}{8}$ inch glass tubing is cut about 5 inches long and bent nearly in the middle at an angle of 120° (see **37**); another piece, 10 inches long, is drawn off at one end to a fine point, and at about three inches from the point is bent at an angle of 60° . The ends are rounded, and a cork is pierced with two holes (see **15**) to receive the tubes (the longest arm of the ten inch tube is forced through the hole in the cork). The wash-bottle is made by fitting the cork furnished with the tubes to an 8 oz. flask. When the bottle is filled with water, by blowing in the short tube, a jet of water can be forced through the longer one, which should plunge nearly to the bottom of the flask. The size of the jet can be regulated by cutting the tapering point at the proper place with a file.

A caoutchouc stopper is preferable to a common one.

10. WATCH-GLASSES. A pair of common watch-glasses are useful for performing some reactions. They can, however, usually be advantageously replaced by bits of broken thin glass-flasks.

APPARATUS.

11. BLOWPIPE. A common jeweler's blowpipe may be used; but a blowpipe with a solid platinum point, such as may be purchased of dealers in chemical apparatus, is preferable.

12. CORKS. One dozen best selected velvet corks of a size to fit the glass flasks. (See **3**)

13. CRUCIBLES. A crucible of Berlin porcelain about an inch in diameter. A porcelain crucible should always be heated by at first waving the lamp flame under it, in order to bring every part gradually and equally to the required tempera-

ture. After this is once effected the strongest heat may be applied without damage. A similar precaution should be observed in cooling it.

14. EVAPORATING DISHES. 2—3 evaporating dishes of Berlin porcelain, 2—4 in. diameter. These dishes also should be heated with great care. A liquid should never be evaporated to dryness in them over the naked flame of a lamp, but the dishes should be placed in a sand-bath or on wire-gauze.

15. FILES. A small round file ($\frac{1}{8}$ in. diameter) and a triangular file are required; the first for boring holes in corks, and the second for cutting glass-tubes.

16. FILTERS.* A few sheets of filter-paper should be cut with scissors into round pieces, using disks of tin or zinc, 2 and 3 in. in diameter, as patterns. A number of these paper disks should be prepared for use. A filter is made by folding a paper disk, so that it takes the form of half of a circle, and folding it again, so that it has the shape of a quarter of a circle. One thickness of the paper thus folded can be separated from the other three, so that a cone of paper is obtained, which fits the shape of a funnel having an angle of 60° . (See 4)

17. FILTER STAND. A wooden apparatus for holding funnels during filtration, which permits of raising or lowering them, can be purchased of dealers in chemical apparatus. The funnel can also be inserted in the neck of the flask, bottle, or test-tube in which the liquid filtered is received. A convenient plan is to have a shelf one inch below the lowest shelf containing the reagents, and to insert a strip of wood, bored with holes, to fit the necks of funnels, in the space between the two shelves. A funnel can thus be held over any part of the desk by placing the strip of wood in the required position.

* The best filter-paper is *soft* Swedish paper, which filters more rapidly than any other, when it is of good quality, and not too hard pressed. In purchasing filtering paper, different samples should always be compared, to test the rapidity of filtration to be obtained with them.

18. FORCEPS. 1 pair of small iron forceps, such as can be purchased at hardware shops.

19. LAMPS. A common alcohol and a Berzelius' alcohol-lamp, or a Bunsen's gas-lamp. The Bunsen's lamp should be provided with a moveable ring to regulate the draught, and the draught should always be adjusted to the height of the flame. (See also 32)

20. MORTAR. A glazed porcelain mortar and pestle.

21. PLATINUM.* A strip of platinum foil, 2 in. by 1 in. weighing about $1\frac{1}{2}$ grammes. A piece of platinum wire, 3 in. long, weighing about $\frac{1}{4}$ gramme, and a piece 3 in. long, weighing about 1-10 grammé.

22. SAND-BATH. A small wrought-iron dish, 4 in. diameter, with a rounding bottom, filled with sand.

23. SPATULA. A thin piece of wood (easily made by the student), or a piece of bone, or a spade-shaped instrument of porcelain, or the knife-blade, is used as a spatula for taking a portion of a precipitate from a filter, or for mixing powders.

24. TRIANGLE. A triangle is made by twisting together the ends of three pieces of iron wire. It should be long enough to rest on the tripod. It is used as a support for the crucible.

25. TRIPOD. A triangle with legs, of $\frac{1}{4}$ in. iron wire, is easily made by a blacksmith. The triangle should be about 5 in. long on each side, and the height of the legs should be adjusted to that of the lamp used.

26. WIRE GAUZE. A piece of brass wire gauze, of stout wire, and with the meshes not more than $\frac{1}{8}$ in. wide, large enough to lay on the tripod. It serves as a support to the vessels heated, and as a protection from the direct action of the flame.

* The higher metals, such as lead, copper, silver, etc., alloy with platinum, and the foil and wire are destroyed when they are heated with substances containing these metals in the pure state, or with substances from which these metals separate on heating, such as their oxides mixed with charcoal or organic matter.

CHEMICAL OPERATIONS.

27. REACTION WITH TEST PAPER. A small piece of red or blue litmus paper is dipped in the solution. If it is acid, blue paper is turned red; if it is alkaline, red paper is turned blue.

Tumeric paper is turned brown by alkalis.

28. PRECIPITATION. When an insoluble body is formed in a solution and separates (falls) from it, precipitation is said to take place. Precipitates are *gelatinous*, as aluminic hydrate; flocculent (consisting of flakes), as sulphide of zinc; or pulverulent, as baric sulphate. Usually the particles are less finely divided, and filtration is easier with precipitates which form in dilute solutions, particularly in boiling solutions. With some precipitates, as magnesian phosphate, in very dilute solutions the act of precipitation is a slow process of crystallization, and the formation of a precipitate does not take place until after several hours.

29. FILTRATION is a process by which an insoluble body, usually a precipitate, is separated from a liquid. It is usually important to allow a precipitate to settle before filtration; and frequently after the clear liquid has been poured upon the filter, it is best to add more water to the precipitate, and to wait again until it has settled. When the precipitate requires to be washed, this process may be repeated many times before the precipitate is brought upon the filter. The precipitate usually clogs the pores of the filter, so that it retards the flow of the liquid. A precipitate which has to be washed is finally brought on the filter by rinsing with the wash-bottle, and it is washed by repeatedly filling the filter with water and allowing it to empty itself. The process should be intermittent, and the filter should never be kept constantly full during the latter part of a filtration, when the object is to wash a precipitate with pure water. Sometimes precipitates, which take the form of a powder, are so finely divided, as to pass through the pores of a filter. This can usually

be avoided by precipitating in a dilute solution, and particularly by boiling the solution. Sulphur cannot be prevented from going through the filter. Filtration is more rapid with hot than with cold water. Before commencing a filtration, the filter should be made to fit closely to the side of the funnel, and it should always be moistened with pure water.

30. DECAN TATION consists in allowing a precipitate to settle, and in pouring off the liquid above it, in the manner already mentioned under filtration. This process may or may not be united with that of filtration. For instance, argentic chloride settles so completely and quickly that it can usually be washed simply by repeated decantations.

31. EVAPORATION is usually performed in a porcelain dish. A few drops of liquid can be evaporated by heating them on the platinum foil. Bits of broken glass or porcelain vessels are very useful for the same purpose.

32. THE USE OF THE BLOWPIPE. An olive oil or kerosene lamp, with a wick $\frac{5}{8}$ in. long and $\frac{1}{4}$ in. broad, or a Bunsen's lamp, with the regulator turned to shut off the draught of air may be used. If the Bunsen's lamp has no regulator for the draught, a smaller tube may be introduced into the lamp-tube, until it rests upon the piece from which the gas issues, and excludes the air. After considerable practice a continuous stream of air can be forced through the blowpipe by making the cavity of the mouth the reservoir, into which air is forced at intervals from the lungs, and is prevented from escaping by a peculiar contraction of the throat and hanging palate, which is easily learned; the breathing goes on through the nose uninterruptedly. The chief difficulty that beginners usually experience, is fatigue of the muscles of the cheeks, which prevents a long-continued effort. Two kinds of flames can be produced with the blowpipe—one containing an excess of air, consequently of oxygen; the other containing an excess of combustible gases, consequently of gases capable of consuming oxygen or reducing.

THE OXYDIZING FLAME is produced by introducing

the point of the blowpipe two-thirds through the lamp-flame. It is a *clear blue* cone, surrounded and continued at the point by a colorless flame, intensely hot, and capable of producing oxydation. The substance should be heated at or just beyond the point of the blue cone.

THE REDUCING FLAME is produced by holding the point of the blowpipe at the outside of the lamp-flame, and by blowing somewhat more gently. The flame is much less pointed and is more luminous than the oxydizing flame. The substance should be heated at a distance from the point of the flame equal to one-third of its length, and should be completely enveloped in the flame. The position of the blowpipe and the force of the blast regulate the quality of the flame. It is a difficult matter to produce the true reducing flame, which should not deposit carbon on the substance heated in it, and at the same time should contain no excess of oxygen.

33. THE BORAX BEAD is formed by making a loop $\frac{3}{8}$ in. in diameter in the stout platinum wire, heating it red hot in the blow-pipe flame, and touching it to a small piece of borax while it is hot. The borax, which adheres to the hot wire, is heated in the blow-pipe flame. It at first swells, while losing its water of crystallization, and finally it melts to a clear glass bead. A finely divided substance can be taken up by touching the hot bead to it, and it can then be tested as to its solubility in the borax bead, coloring properties, &c., in the blow-pipe flame.

34. HEATING ON CHARCOAL. Select a good piece of charcoal, at least 4 in. long and 1 in. broad and thick, and smooth a plane surface in a direction at a right angle with that of the year-rings. (If heat is applied to a surface parallel to the planes of the year-rings, the charcoal is more liable to snap from the expulsion of moisture.) In many cases the charcoal serves as a convenient support for a substance to be heated; in others the reducing agency of the charcoal comes in play. Substances are also evaporated at a high temperature from the surface of the charcoal.

35. DUCTILITY, MALLEABILITY, BRITTLE-NESS are characteristic properties of metals, and metals can be tested with regard to them by pounding with a hammer or by rubbing with the pestle of a mortar. When a substance which appears to contain a higher metal is reduced by sodic carbonate on charcoal, unless metallic globules are at once apparent, the portion of the charcoal which has been heated should be cut out and pulverized with water in a mortar, and washed by decantation. If metallic globules have been formed, they will sink to the bottom, and after thorough washing, during which the sodic carbonate is dissolved, and the light particles of charcoal are floated away, they will appear as globules, if they are hard like copper; as brittle grains, if they are brittle like bismuth; or as flattened disks, when they are ductile like lead, and when they have been pressed by the pestle against the mortar.

36. COLOR OF THE FLAME. If the substance to be tested is a solid, a small piece of it is brought on a loop of the fine platinum wire, or in a pair of forceps, into the flame of the alcohol or Bunsen's lamp, and the color imparted to the flame observed. If a substance in solution is to be tested, the fine platinum wire is dipped in the solution and is then introduced into the lamp-flame. If the solution is too dilute to afford a distinct test in this way, it must be evaporated, and it is usually best to evaporate nearly to dryness, and to take some of the solid residue for the test.

37. THE MANIPULATION OF GLASS TUBING. Glass softens, when a small piece is heated in the flame of an alcohol lamp, or when a larger piece is heated in a Bunsen's lamp, or with the blowpipe flame or in a blast lamp.

A tube can be bent easily, as soon as the glass softens.

It is best only to bend gently at first, then to heat the adjacent part of the tube, and to bend again, and so on, in order that the sides of the tube may not fall together in bending.

When a glass tube is heated for some time, it contracts and the sides thicken. By drawing out a tube either immediately

after it has become soft, or after the sides have thickened, a tapering point of any desired calibre and thickness of glass can be obtained.

To close the end of a glass tube, draw the tube off while the glass is as thin as possible, and hold the tapering point in the flame, and draw the end off again; in this way a tube with a pointed, closed end is obtained; by heating the closed end of the tube removing it from the flame, rotating it and blowing in it, while the glass is still red-hot, it expands, and a more rounded end or a bulb can be produced. Glass tubing can be cut by making a mark at the required place by a few file-strokes, and then by breaking the glass. The ends of glass tubes cut in this way should be held in the flame till they become red-hot; in this way the sharp edges become rounded.

PART II.

PART II is preparatory to Part III, which contains a general scheme of analysis applicable to compounds of all the elements. The most important tests are those which are described in Part III, and a knowledge of them would suffice alone for the purposes of analysis, if the liability to error in chemical manipulations did not make it expedient to employ a variety of tests, as corroborative evidence, before coming to a conclusion in regard to the composition of a substance.

It is important that the student should turn to Part III, and commit to memory the general features of the scheme of separation for each group, at the time that he is performing the reactions of the members of the group as they are described in the following pages. By this means he will make himself familiar with the important points in which the compounds with which he has to deal differ from each other, and the manner in which these differences can be used in analysis; also at this stage of his progress it is advisable for him to make mixtures of compounds of several elements, and to analyze them according to the directions given in Part III.

The student, keeping in view the reasons for learning the characteristic reactions of the compounds of each element, should perform carefully the tests described in Part II, supplementing the description by the closest observation of the phenomena as they pass before his eyes. By practice of this kind he will soon acquire the skill in manipulation necessary for analytical work. Always, when a reaction is performed, the equation describing it should be written. The formulas of the reagents and the com-

pound operated upon * form the first half of the equation; the formula of the precipitate, which is given in the book, enters into the second half of the equation and determines the formulas of its other members.

Thus it is known that baric chloride and calcic sulphate give a precipitate of BARIC SULPHATE, BaSO_4 . (See page 32.) From the formulas on the labels of the bottles we can construct the equation: $\text{BaCl}_2 + \text{CaSO}_4 = \text{BaSO}_4 + \text{X}$, and by inspecting the equation we find the unknown quantity X can only be CaCl_2 . The following case is more complicated: disodic hydric phosphate, ammoniac hydrate, and magnesian sulphate, form a precipitate of MAGNESIAN AMMONIAC PHOSPHATE, MgNH_4PO_4 ; or, putting the statement into formulas, $\text{Na}_2\text{HPO}_4 + \text{NH}_4\text{HO} + \text{MgSO}_4 = \text{MgNH}_4\text{PO}_4 + \text{X}$. Here $\text{X} = \text{Na}_2 + \text{H}_2 + \text{O} + \text{SO}_4$, and the question arises: How are these bodies combined? A slight experience will teach that the rule 2d (page 8) brings the SO_4 and the Na together, and consequently the H_2 and the O; while an inspection of the tables of atomicities (page 2 and page 11) shows that SO_4 combines with 2 Na, and that O combines with 2 H; hence X becomes $\text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$.

The grouping together of the elements or groups of elements appearing in reactions is not usually a difficult matter, and is soon learned with practice.

The formulas of the compounds which are most frequently used in the laboratory stand after the names of the metals and acids, and can be used in writing equations.

* These formulas should be given in full upon the labels of the bottles containing the compounds used and the reagents.

TESTS FOR METALS.

GROUP I.

SODIUM, POTASSIUM AND AMMONIUM.

THERE is no reagent which precipitates all the metals of this group. The salts of metals of Group I have a neutral reaction when they contain strong acids like chlorhydric, nitric, and sulphuric acids. They have an alkaline reaction when they contain weak acids like sulphydric, boracic and carbonic acids.

SODIUM.



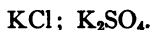
Sodium compounds can be recognized by heating them in the loop of a piece of fine platinum wire (36) in the flame of a lamp. Sodium colors the flame *yellow*, and can be recognized, even when mixed with much larger quantities of other elements, which alone impart other colors to the flame.

When a liquid is to be tested, it may be evaporated and the residue brought on the platinum wire, or frequently it is sufficient to dip the wire in the liquid and to bring it into the flame of the lamp.

No reagent* precipitates sodium compounds.

* Only the reagents spoken of in this book are referred to.

POTASSIUM.



Potassium compounds color the flame of a lamp *violet*. A small admixture of sodium obscures the color of the flame of potassium, but the sodium color disappears, and that characteristic of potassium can be observed, when the flame is viewed through a thick glass, colored blue with cobalt.

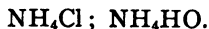
Mix together a sodium and a potassium salt, observe the yellow color imparted to the flame by the mixture, showing the presence of sodium, and then examine the flame for the potassium color through a piece of cobalt glass thick enough to exclude the sodium flame. Examine a pure sodium flame with cobalt glass to ascertain that the glass does not allow the color of the sodium to pass through it, or until the blue color of the sodium flame can be easily distinguished from the violet of the potassium flame.*

PLATINIC CHLORIDE precipitates concentrated solutions of potassic chloride, as a DOUBLE CHLORIDE OF PLATINUM and POTASSIUM, $(\text{KCl})_2 \text{PtCl}_4$. No other salt of potassium can be used for this test. It is best to evaporate the solution to dryness in a water-bath, with a large quantity of platinic chloride, and to wash the residue several times with alcohol. The double chloride is left as a *yellow crystalline powder*, which gives the potassium flame. The double chloride is slightly soluble in water but insoluble in alcohol.

No other reagent precipitates potassium compounds.

* If the glass is sufficiently thick and intense in color, the light from the sodium flame is completely excluded. A thinner glass allows a part of the light to pass through, but it then has a blue color which can be distinguished after practice from the violet color of the potassium flame, which passes through the glass with little of its brilliancy diminished.

AMMONIUM.



Ammonium compounds do not color the flame of a lamp.

PLATINIC CHLORIDE precipitates ammoniac chloride as a DOUBLE CHLORIDE OF PLATINUM AND AMMONIUM $(\text{NH}_4\text{Cl})_2\text{PtCl}_4$. The precipitate forms under the same circumstances, and has the same aspect and properties as that obtained with potassic chloride, but it can be distinguished from the latter by the absence of color imparted to the flame of a lamp, when it is heated in it. It is destroyed at a dull red heat and a residue of metallic platinum is left.

SODIC HYDRATE, added in excess to ammoniac compounds, causes them to give off the smell of AMMONIA GAS, NH_3 , especially when the solution is heated. AMMONIA GAS colors moist tumeric paper *brown* (a delicate test).

No other compound interferes with the application of this test.

GROUP II.

BARIUM, CALCIUM AND MAGNESIUM.

The chlorides and nitrates of metals of Group II are soluble in water. The sulphates of calcium and magnesium are also soluble. The solutions have a neutral reaction with test paper.

AMMONIC and SODIC CARBONATES precipitate the metals of Group II in neutral solutions as CARBONATES. The carbonate of magnesium is very soluble in solutions of ammoniac salts, particularly in ammoniac chloride; therefore no precipitate of magnesian carbonate is produced when these salts are present in considerable quantity.

NEUTRAL PHOSPHATES (as disodium hydric phosphate) precipitate the metals of Group II in neutral or alkaline solutions as PHOSPHATES.

The carbonates and phosphates of metals of Group II are soluble in dilute acids, unless the acids themselves are capable of precipitating the metals.

SULPHYDRIC ACID, AMMONIC SULPHIDE, and AMMONIC HYDRATE, do not precipitate the metals of Group II.

BARIUM.



SULPHURIC ACID (dilute) precipitates baric compounds, as BARIC SULPHATE, BaSO_4 , *white powder*.

CALCIUM SULPHATE and other soluble sulphates give the same precipitate with baric compounds.

Baric sulphate is insoluble in acids.

Barium compounds, particularly when moistened with chlorhydric acid, color the flame *yellowish green*.

CALCIUM.

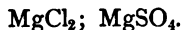


SULPHURIC ACID (dilute) and soluble sulphates, except calcic sulphate, precipitate concentrated solutions of calcium compounds, as CALCIC SULPHATE, CaSO_4 , *white powder*. Calcic sulphate is soluble in a considerable quantity of water, therefore no precipitate is produced in very dilute solutions by sulphuric acid. It is insoluble in dilute alcohol, therefore sulphuric acid (dilute), with the addition of a large quantity of alcohol, precipitates calcic sulphate from even dilute solutions of calcium compounds.

AMMONIC OXALATE precipitates calcic compounds from neutral or alkaline solutions, as CALCIC OXALATE, CaC_2O_4 , *white powder*. The precipitate forms best after standing some time in a solution to which ammonic hydrate has been added in excess. Calcic oxalate does not dissolve in acetic acid. Baric salts, but not magnesian salts in the presence of ammonic chloride, are precipitated by ammonic oxalate in solutions, to which ammonic hydrate has been added in excess.

Calcium compounds, particularly when moistened with chlorhydric acid, color the flame *yellowish red*.

MAGNESIUM.



DYSODIC HYDRIC PHOSPHATE precipitates magnesian compounds, to whose solution ammonic hydrate and ammonic chloride have been added, as MAGNESIAN AMMONIC PHOSPHATE, MgNH_4PO_4 , *white crystalline powder, or white flakes if the solution is concentrated*. This precipitate only appears after the lapse of some time in very dilute solutions. It is then crystalline.

SODIC HYDRATE, in excess, precipitates magnesian compounds as MAGNESIC HYDRATE, $Mg(HO)_2$, *while powder*, when it is boiled with their solutions. In case ammoniac chloride is present, the boiling must be continued until the odor of ammonia is no longer perceptible.

Mix together the chlorides of barium, calcium and magnesium and dilute the solution with a considerable quantity of water. Add dilute sulphuric acid and boil. If the solution is sufficiently dilute, the *barium* alone is precipitated. Filter a small portion of the solution containing the precipitate and add sulphuric acid to the filtrate; if a precipitate forms, all of the barium has not been precipitated and more sulphuric acid must be added to the remainder of the solution until the barium is completely precipitated. Filter and preserve the barium precipitate on the filter. Add a small quantity of ammoniac chloride, then ammoniac hydrate until the reaction becomes alkaline, and finally ammoniac oxalate. The *calcium* alone is precipitated if the solution is free from barium. Add enough ammoniac oxalate to precipitate all the calcium, testing with ammoniac oxalate in the same way as before in order to ascertain when the precipitation is complete. Filter and preserve the calcium precipitate. Add to the filtrate disodium hydric phosphate. *Magnesium* alone is precipitated, when the solution is free from barium and calcium. Wash the barium and calcium precipitates on the filters until a drop of the water which passes through the filter leaves no residue after evaporation on platinum foil, and test the color which each precipitate imparts to the flame.*

* Processes similar to the above are used for the separation of all the metals from each other, and care must always be taken to ascertain whether enough of a reagent has been added to completely effect a precipitation before the next test is proceeded with in the filtrate, and when a precipitate is to be examined, it must be thoroughly freed by washing from the solution which adheres to it.

GROUP III.

ALUMINIUM AND CHROMIUM.

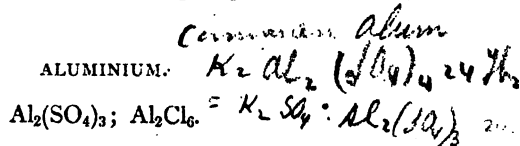
The sulphates, chlorides, and nitrates of metals of Group III are soluble in water, and the solution has an acid reaction with test paper. Aluminic and chromic alum solutions have a neutral reaction.

AMMONIC HYDRATE, CARBONATE, and SULPHIDE precipitate the metals of Group III, as **HYDRATES**.

NEUTRAL PHOSPHATES precipitate the metals of Group III as **PHOSPHATES**.

SULPHYDRIC ACID does not precipitate the metals of Group III.

ALUMINIUM.



AMMONIC HYDRATE precipitates aluminic compounds as **ALUMINIC HYDRATE**, $\text{Al}_2(\text{HO})_6$, *gelatinous, white flakes*.

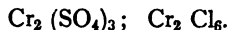
The precipitate is insoluble in ammoniac hydrate and in ammoniac chloride, but dissolves in acids and in sodic hydrate. It forms best on boiling.

SODIC HYDRATE precipitates aluminic compounds like ammoniac hydrate, but an excess of sodic hydrate dissolves the precipitate so quickly that its formation easily escapes notice.

No precipitate is formed when the solution in sodic hydrate is boiled.

Solid compounds of aluminium (except silicates), when moistened with cobaltic nitrate solution, and heated with the oxidizing blowpipe-flame (see page 24, **34**) on charcoal, or on a platinum wire, take a *blue color*.

CHROMIUM.



Solutions of chromic oxide compounds are *green*.

AMMONIC HYDRATE precipitates chromic compounds as CHROMIC HYDRATE, $\text{Cr}_2(\text{HO})_6$, *gelatinous, dirty green flakes*. The precipitate is insoluble in ammoniac hydrate after boiling, and in ammoniac chloride, but dissolves in acids and in sodic hydrate.

SODIC HYDRATE precipitates chromic compounds like ammoniac hydrate, but dissolves them, when an excess of sodic hydrate is present. CHROMIC HYDRATE is precipitated from its solution in sodic hydrate, when the *dilute* solution is boiled for some time.

BLOWPIPE REACTIONS. Compounds of chromium color the borax bead *green*.

If chromic hydrate, or any solid chromic compound, is mixed with equal parts of sodic carbonate and sodic nitrate, and heated red-hot on the platinum foil, chromate of sodium is formed by the oxidation of chromic oxide. Chromate of sodium dissolves in water with a yellow color. The color is heightened when an acid is added, and an acid chromate is formed in the solution.

This reaction is a characteristic test for chromium compounds.

Mix together solutions of chromium and aluminium salts, add sodic hydrate until the reaction becomes very strongly alkaline (the precipitate which first forms will dissolve), dilute with a considerable quantity of water in a small flask, and boil for several minutes after a dirty green precipitate has formed. Filter from the precipitate. This precipitate contains all the *chromium*. Test it according to Part III (99). The filtrate contains all the *aluminium*. Test it according to Part III (101).

GROUP IV.

ZINC, MANGANESE, IRON, NICKEL AND COBALT.

The sulphates, chlorides and nitrates of metals of Group IV are soluble in water. The solutions have an acid reaction with test paper.

AMMONIC SULPHIDE precipitates the metals of Group IV as sulphides; if the solution is not neutral, it should be made so with ammonic hydrate.

SODIC HYDRATE and AMMONIC HYDRATE precipitate the metals of Group IV as HYDRATES. The HYDRATE OF ZINC is soluble in an excess of the precipitant, and the HYDRATES OF NICKEL AND COBALT are soluble in ammonic hydrate.

SODIC CARBONATE precipitates the metals of Group IV as CARBONATES (ferric compounds as ferric hydrate).

NEUTRAL PHOSPHATES precipitate the metals of Group IV as PHOSPHATES.

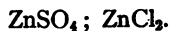
SULPHYDRIC ACID does not precipitate the metals of Group IV, when they are in an acid solution. (See Zinc, page 38.)

SECTION I.

ZINC, MANGANESE AND IRON.

Metals whose sulphides are soluble in cold dilute chlorhydric acid.

ZINC.



Metallic zinc dissolves readily in dilute chlorhydric acid and sulphuric acid, with evolution of hydrogen.

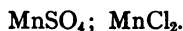
The metal melts readily when heated on charcoal with the blowpipe, and at a high temperature it distils, and the vapor burns with a *bluish-white* flame, depositing an incrustation on the charcoal of OXIDE OF ZINC, ZnO , which is *white* when hot, and *yellow* when cold. If the incrustation is moistened with cobaltic nitrate and heated in the oxydizing flame, it turns *dirty green*. By this test zinc can often be recognized in alloys.

AMMONIC SULPHIDE precipitates zinc compounds as the SULPHIDE OF ZINC, ZnS —*white, flocculent precipitate*. The sulphide of zinc is soluble in dilute chlorhydric acid, but not in acetic acid. It is the only white insoluble sulphide.

SULPHYDRIC ACID precipitates zinc as the SULPHIDE OF ZINC, when the metal is only combined with acetic acid. To obtain the precipitate, if a stronger acid is present, add sodic hydrate until the solution has a strongly alkaline reaction, and then add acetic acid, until the reaction becomes acid, before treating with sulphydric acid.

SODIC and AMMONIC HYDRATES precipitate zinc compounds as the HYDRATE OF ZINC, $\text{Zn}(\text{HO})_2$, *white flakes*. The precipitate dissolves easily in an excess of the precipitants; and the solution in sodic hydrate is not reprecipitated when it is boiled.

MANGANESE.



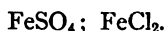
Solutions of manganese compounds are of a *faint pink* color.

AMMONIC SULPHIDE precipitates manganese compounds as the SULPHIDE OF MANGANESE MnS , *flesh-colored flakes*. The sulphide of manganese is soluble in dilute acids.

SODIC and AMMONIC HYDRATES precipitate manganese compounds, as MANGANOUS HYDRATE, $\text{Mn}(\text{HO})_2$, *white flakes, which turn brown on exposure to the air*. Manganous hydrate is insoluble in an excess of the precipitant, but it is soluble in a large quantity of ammonic chloride.

BLOWPIPE REACTIONS. Manganic compounds color the borax bead *amethyst* in the oxydizing flame. If a compound of manganese is heated with a soda bead (which can be made in the same way as a borax bead in the loop of a platinum wire) in the oxydizing flame, it colors it *green*, in consequence of the formation of manganate of sodium. The same color is produced, when a compound of manganese is heated on the platinum foil with sodic carbonate and nitrate. On boiling the green salt with water, it is destroyed, the color disappears, and brown flakes of manganic hydrate are precipitated.

IRON (*ferrous salts*).



Metallic iron dissolves readily in dilute acids, with evolution of hydrogen.

Ferrous salts in solution have a *pale green* color.

OXYDIZING AGENTS (as nitric acid and potassic chlorate), when heated with acid solutions of ferrous salts, oxydize them to **FERRIC SALTS**, whose color is *brownish red*, or *reddish yellow*, and is more intense than the *green* color of ferrous salts.

AMMONIC SULPHIDE precipitates ferrous salts, as **FERROUS SULPHIDE**, FeS , *black flakes*. Ferrous sulphide is soluble in dilute acids.

SODIC and AMMONIC HYDRATES precipitate ferrous compounds as **FERROUS HYDRATE**, $\text{Fe}(\text{HO})_2$, *at first nearly white, then bluish green, and, finally, by absorption of oxygen, reddish brown*. Ferrous hydrate is insoluble in a excess of sodic hydrate. The presence of a large quantity of ammonic salts in a solution prevents its precipitation.

POTASSIC FERROCYANIDE precipitates ferrous compounds, as **POTASSIC FERROUS FERROCYANIDE**, $\text{K}_2\text{Fe}(\text{FeCy}_6)$, *

* Cy, cyanogen, is used as a symbol for the group CN.

bluish white, turning quickly *dark blue*, through absorption of oxygen from the air.

POTASSIC FERRICYANIDE precipitates ferrous compounds, as TURNBULL'S BLUE, $\text{Fe}_3(\text{Fe}_2\text{Cy}_{12})$,* *deep blue*. This is the best test for ferrous compounds.

The last two precipitates are insoluble in dilute acids.

POTASSIC SULPHOCYANATE gives no coloration with ferrous compounds.

IRON (*ferric salts*). *See Fe, Fe₂, Fe₃, Cl₃*

$\text{Fe}_2\text{Cl}_6 =$ *See Fe₂, Fe₃, Cl₃*

REDUCING AGENTS, as sulphurous and sulphydric acids, metallic zinc and iron, reduce ferric salts in solution to ferrous salts, when a free acid is present.

The reaction with sulphydric acid is accompanied by a precipitation of sulphur $\text{Fe}_2\text{Cl}_6 + \text{H}_2\text{S} = 2\text{FeCl}_2 + 2\text{HCl} + \text{S}$.

A similar reaction takes place with ammoniac sulphide and the other salts of sulphydric acid.

Ferric salts in solution have a *yellow* color, and they possess a much stronger coloring power than ferrous salts.

AMMONIC SULPHIDE reduces ferric salts to ferrous salts, and then precipitates FERROUS SULPHIDE, FeS , *black flakes*.

SODIC and AMMONIC HYDRATES precipitate ferric compounds as FERRIC HYDRATE, $\text{Fe}_2(\text{HO})_6$, *red gelatinous flakes*, insoluble in an excess of the precipitants, and in ammoniac salts.

POTASSIC FERROCYANIDE produces a precipitate of PRUSSIAN BLUE, $\text{Fe}_4(\text{FeCy}_6)_3$, *deep blue*, in a solution of ferric salts.

POTASSIC FERRICYANIDE colors solutions of ferric salts *deep reddish brown*, but produces no precipitate. On the addition of a reducing agent a *deep blue* precipitate forms.

* See note on preceding page.

POTASSIC SULPHOCYANATE gives a *deep red color*,* with the smallest traces of ferric compounds in acid solutions.

The different behavior, with the last three reagents of ferrous and ferric salts serves to distinguish between them.

BLOWPIPE REACTION. Iron colors the borax bead *green* in the reducing flame, and *reddish yellow while hot*, and *yellow while cold*, in the oxydizing flame.

Mix together solutions of zinc, manganese and iron salts; if a ferrous salt is taken, add a little chlorhydric acid and boil the solution for a few minutes with one or two crystals of potassic chlorate, in order to convert the ferrous into ferric salt. Add sodic hydrate to the solution until the reaction is strongly alkaline, boil for a few minutes and filter. All the *manganese* and *iron* will be precipitated. Test the precipitate for iron according to Part III (97), and for manganese according to Part III (98). All the *zinc* will be contained in the filtrate; test it for *zinc* according to Part III (100).

SECTION II.

NICKEL AND COBALT.

Metals whose sulphides are insoluble in cold, dilute chlorhydric acid.

NICKEL.



Solutions of nickel salts are *green*.

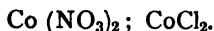
AMMONIC SULPHIDE precipitates nickel compounds as the SULPHIDE OF NICKEL, NiS , *black flakes*. The sulphide of nickel is insoluble in cold, dilute chlorhydric acid. It dissolves readily on boiling, or in a strong acid.

* Potassic sulphocyanate gives the same color with a solution containing a large quantity of free nitric acid.

SODIC and AMMONIC HYDRATES precipitate nickel compounds as the HYDRATE OF NICKEL, $\text{Ni}(\text{HO})_2$, *apple green*. The hydrate of nickel is insoluble in an excess of sodic hydrate. It dissolves in ammoniac hydrate, and the solution has a *blue* color.

BLOWPIPE REACTIONS. Nickel colors the borax bead in the oxydizing flame *violet, when it is hot, and a faint reddish-brown, when it is cold*. The latter color can best be seen by flattening the bead when it is hot, by pressure, with the end of the blowpipe on a glass surface. By long-continued reduction in the reducing flame, or on charcoal, the bead may be obtained *colorless*, but with gray specks of reduced metal in it.

COBALT.



Solutions of cobalt salts, when dilute, are *red*.

AMMONIC SULPHIDE precipitates cobalt compounds as the SULPHIDE OF COBALT, CoS , *black flakes*. The sulphide of cobalt is insoluble in cold dilute chlorhydric acid. It dissolves readily on boiling or in a strong acid.

SODIC and AMMONIC HYDRATES precipitate cobalt compounds at first as a *blue basic salt*, which changes to the *pale red* COBALTOUS HYDRATE, $\text{Co}(\text{HO})_2$ on boiling. On exposure to the air it becomes brown, through the formation of cobaltic hydrate. Cobaltous hydrate is insoluble in an excess of sodic hydrate, but it dissolves in ammoniac hydrate, and the solution is *red, tinged with brown*.

BLOWPIPE REACTIONS. Cobalt compounds color the borax bead *blue*, and the color is so intense that a small quantity of cobalt eclipses the color produced by a much larger quantity of nickel, when the latter is mixed with it. The blue color does not disappear on reduction, so that when sufficient nickel is present to hide the color of a small quantity of cobalt in a bead, the color of the nickel may be made to disappear by a thorough reduction, either on the platinum wire, or on charcoal,

so that the blue color characteristic of cobalt can be detected in the bead, after it has been melted and flattened by pressure with the end of the blowpipe on a glass surface. If the bead was reduced on charcoal, it is advisable to remove it from the charcoal and to melt it on the platinum wire in order to examine its color.

Mix together sulphates of zinc, manganese and nickel and the nitrate of cobalt and ferrous sulphate, add ammoniac hydrate until a permanent precipitate begins to form, and then ammoniac sulphide* until the metals are completely precipitated as sulphides. Wash the precipitate on a filter and treat it with cold, dilute chlorhydric acid in order to separate the sulphides of *nickel* and *cobalt* from the other sulphides. See Part III (92).

Test for nickel and cobalt according to Part III (93) and (94).

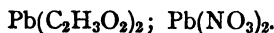
Test for the other metals in the chlorhydric acid solution according to Part III (95), (96), (97), (98) and (100).

* In order to precipitate the nickel completely, the solution must not contain free ammoniac hydrate and the ammoniac sulphide must not contain an excess of ammoniac hydrate. If these precautions are not observed a part of the sulphide of nickel dissolves, imparting a black color to the solution.

CHLORHYDRIC ACID precipitates mercurous compounds as MERCUROUS CHLORIDE, Hg_2Cl_2 , *white powder*. Mercurous chloride turns black, but does not dissolve when ammoniac hydrate is added to it. It is insoluble in dilute acids. It dissolves in aqua regia.

SODIC and AMMONIAC HYDRATES give with mercurous compounds *black* precipitates, insoluble in an excess of the precipitant. (For tests by heating mercurous compounds, see page 48.)

LEAD.



Metallic lead dissolves readily in nitric acid.

SULPHURIC ACID and AMMONIAC SULPHIDE precipitate lead compounds, as the SULPHIDE OF LEAD, PbS , *black*. Sulphide of lead is oxidized by strong nitric acid, with formation of sulphate of lead, *white powder*, which is insoluble, unless a very large quantity of nitric acid is present.

CHLORHYDRIC ACID precipitates lead compounds, as the CHLORIDE OF LEAD, PbCl_2 , *white*. When the solution is dilute no precipitate is produced. The chloride of lead dissolves entirely on boiling with a large quantity of water.

SULPHURIC ACID precipitates lead compounds, as the SULPHATE OF LEAD, PbSO_4 , *white powder*. Sulphate of lead is soluble to some extent in chlorhydric and nitric acids, and it is slightly soluble in water. It is insoluble in a mixture of alcohol and water. When complete precipitation is required, it is best to evaporate with an excess of sulphuric acid, until all the other acids are driven off, then to dilute with water, and to add an equal bulk of alcohol.

LEAD, BARIUM, CALCIUM and STANNIC COMPOUNDS are the *only ones precipitated* by SULPHURIC ACID.

SODIC and AMMONIAC HYDRATES precipitate compounds of lead, as BASIC SALTS OF LEAD, *white precipitate*. Basic salts of lead are somewhat soluble in sodic hydrate.

BLOWPIPE REACTIONS. Solid compounds of lead give, when heated with sodic carbonate on charcoal, globules of metallic lead, recognizable by their softness and ductility. An incrustation of OXIDE OF LEAD, PbO , *deep yellow when hot, light yellow when cold*, is formed upon the charcoal, not far from the place where the substance is heated.

Pure lead, or alloys containing a large amount of lead, when heated on charcoal, without soda, burn with a *blue* flame, and give the PbO incrustation.

Mix together acetate of lead, nitrate of silver and mercurous nitrate, add dilute chlorhydric acid until a precipitate ceases to form on further addition and filter the liquid. The precipitate contains all the *silver* and *mercury* and a part of the *lead* as chlorides. Add to the filtrate an equal bulk of alcohol and a small quantity of dilute sulphuric acid. All the lead which it contains will be precipitated. Filter, wash the precipitate and test it for lead on charcoal.

Make a hole in the bottom of the filter and wash the precipitate obtained with chlorhydric acid into a small flask. Test the precipitate for the remainder of the lead, and for silver and mercury according to Part III (63), (64) and (65).

SECTION II.

MERCURIC COMPOUNDS, COPPER AND BISMUTH.

Metals which are not precipitated by Chlorhydric Acid.

MERCURIC COMPOUNDS.



SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate mercuric compounds, at first as *double salts*, which appear first *white*, then *yellow*, *orange*, and *brown*, and finally as MERCURIC SULPHIDE, HgS , *black*. Mercuric sulphide does not

dissolve when it is boiled with moderately concentrated nitric or chlorhydric acid. It dissolves readily in aqua regia.

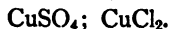
SODIC HYDRATE precipitates mercuric compounds at first as *basic salts, reddish brown*, finally as MERCURIC OXIDE, HgO , *yellow*; insoluble in an excess of the precipitant.

AMMONIC HYDRATE precipitates mercuric compounds, as SALTS CONTAINING AMMONIA, *white*. The precipitate is insoluble in an excess of ammoniac hydrate.

STANNOUS CHLORIDE reduces mercuric compounds, and precipitates them as MERCUROUS CHLORIDE, Hg_2Cl_2 , *white*. After the metals of Group V, Section I, if they are present, have been removed from a solution by the addition of chlorhydric acid, mercuric compounds are the only ones which give a precipitate in acid solution with stannous chloride.

REACTIONS WITH THE AID OF HEAT. Dry mercurous and mercuric chlorides form white sublimes, when they are heated in a closed glass tube. All dry compounds of mercury, when they are mixed with dry sodic carbonate, and heated in a closed tube, give a sublimate of metallic mercury. The sublimate is at first a faint metallic film, which augments, until drops of mercury appear. If the quantity of mercury is small, the film may be made to take the form of metallic drops by rubbing it with a copper wire.

COPPER.



Metallic copper dissolves readily in dilute nitric acid. It dissolves with difficulty in chlorhydric acid. All copper solutions are *blue*.

IRON or ZINC precipitates copper from its acid solutions, either as a metallic coating or as *brownish red* metallic grains. If a strip of zinc and one of platinum are placed in a dilute acid solution of copper, so that they touch each other, the platinum is plated with copper.

SULPHYDRIC ACID and **AMMONIC SULPHIDE** precipitate cupric compounds, as CUPRIC SULPHIDE, CuS , *black*. Cupric sulphide is insoluble in dilute acids, but dissolves readily in strong acids. It is somewhat soluble in an excess of ammoniac sulphide.

SODIC HYDRATE precipitates cupric compounds as CUPRIC HYDRATE, $\text{Cu}(\text{HO})_2$, *light blue*. On boiling, CUPRIC OXIDE, CuO , *black*, is formed. The precipitate is insoluble in an excess of sodic hydrate.

AMMONIC HYDRATE precipitates cupric compounds, as CUPRIC HYDRATE, which dissolves immediately in an excess of ammoniac hydrate. The solution has a very intense *blue* color.

A valuable test for cupric compounds.

FERROCYANIDE OF POTASSIUM precipitates cupric compounds in acid solutions, as FERROCYANIDE OF COPPER, $\text{Cu}(\text{FeCy}_6)$, *reddish brown*. This is a delicate test for very small quantities of copper.

BLOWPIPE REACTIONS. Cupric compounds color the borax bead *blue* when cold, and *green* when hot, in the oxydizing flame. The bead is colored *red* and becomes *opaque* in the reducing flame. No other metal produces this coloration.

When cupric compounds are heated with sodic carbonate on charcoal, metallic copper is reduced in the form of small globules, which can be easily recognized by their *hardness and red color*.

COPPER FLAME. Compounds containing copper (alloys and salts) color the flame of a lamp *green*, or if chlorine is present, *blue*. By moistening a cupric compound with chlorhydric acid the *blue color* can easily be detected.

BISMUTH.



Metallic bismuth dissolves readily in moderately concentrated

nitric acid. It dissolves with great difficulty in chlorhydric acid.

WATER. Solutions of bismuth, particularly those containing chlorhydric acid, are remarkable for giving a precipitate consisting of a BASIC SALT, when water is added to them, unless they contain a large quantity of free acid. The basic salt can be dissolved by the addition of an acid. CHLORHYDRIC ACID precipitates nitrate of bismuth solution as a BASIC CHLORIDE because the basic chloride of bismuth is more insoluble than the other basic salts. The precipitate is soluble on the further addition of chlorhydric acid.

SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate bismuth compounds as the SULPHIDE OF BISMUTH, Bi_2S_3 , *black*.

SODIC and AMMONIC HYDRATES precipitate bismuth compounds as the HYDRATE OF BISMUTH, $\text{Bi}(\text{HO})_3$, *white*.

The hydrate of bismuth is insoluble in an excess of sodic and ammonic hydrates.

BLOWPIPE REACTIONS. Bismuth compounds, mixed with sodic carbonate, and heated on charcoal, give *brittle* metallic globules, and an incrustation of OXIDE OF BISMUTH, Bi_2O_3 , on the charcoal.

The incrustation is *orange yellow* when hot, and *bright yellow* when cold.

Mix together mercuric chloride, cupric sulphate and nitrate of bismuth, and add sulphydric acid until the metals are completely precipitated. Filter and wash the precipitate with a little water. Spread the filter out, scrape the precipitate off from it and heat the precipitate gently in a porcelain dish with strong nitric acid until red fumes cease to be given off, then add a little water and boil for a few minutes.

The sulphide of *mercury* remains insoluble. Test it according to Part III (86). Test the filtrate for *bismuth* and *copper* according to Part III (88) and (89).

GROUP VI.

TIN, ANTIMONY, ARSENIC AND GOLD.

The metals of Group VI sometimes act as acids, uniting with metals, and sometimes as metals, uniting with acids. Their combinations with acids have an acid reaction; those with metals which are soluble in water, have an alkaline reaction.

SULPHYDRIC ACID precipitates the metals of Group VI as **SULPHIDES**. The precipitation takes place slowly, particularly in the case of arsenic acid, and it should never be considered complete unless a current of sulphydric acid gas is passed through the solution for some time, and it is left to stand twenty-four hours.

The sulphides of metals of Group VI are insoluble in dilute acids, but they are decomposed or dissolved by boiling with concentrated acids. They are soluble in sodic hydrate and in ammoniac sulphide (the sulphide of gold with difficulty).

TIN.

Sn.

Metallic tin dissolves readily in strong chlorhydric acid on boiling, and the solution contains **STANNOUS CHLORIDE**, SnCl_2 . Tin is oxydized to **STANNIC OXIDE**, SnO_2 , *white powder*, by strong nitric acid. Stannic oxide is insoluble in nitric acid, but dissolves readily in hot concentrated chlorhydric acid, and the solution contains **STANNIC CHLORIDE**, SnCl_4 . Stannic chloride is also formed by the action of aqua regia on metallic tin.

METALLIC ZINC precipitates **METALLIC TIN** from its solutions in acids as crystalline metallic particles.

BLOWPIPE REACTIONS. Compounds of tin are reduced, when they are heated with sodic carbonate and potas-

sic cyanide on charcoal; and METALLIC TIN may be discovered in flattened globules by rubbing the fused mass, taken from the charcoal, in a mortar with water, and washing several times by decantation. When the globules are large, they may be observed on the charcoal during the fusion.

OXYDATION by SODIC NITRATE. When a sulphide of tin is oxydized at the lowest possible temperature by a mixture of sodic nitrate and carbonate, STANNIC OXIDE, SnO_2 , which is insoluble in water, is formed. If the oxydation is carried on at too high a temperature, stannate of sodium is formed, which is soluble in water.

STANNOUS COMPOUNDS.



SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate stannous compounds from acid solutions as STANNOUS SULPHIDE, SnS , *dark brown*. Stannous sulphide dissolves with difficulty in a colorless solution of ammonic sulphide (the mono-sulphide), and it is scarcely soluble in ammonic hydrate and carbonate. It is converted into stannic sulphide by a yellow solution of ammonic sulphide (the poly-sulphide), and dissolves readily when warmed with the solution. Stannous sulphide is soluble in sodic hydrate. The sulphides of tin are precipitated from these solutions, when a dilute acid is added gradually, until the reaction becomes strongly acid.

SODIC and AMMONIC HYDRATES and CARBONATES precipitate stannous compounds as STANNOUS HYDRATE, $\text{Sn}(\text{HO})_2$, *white*.

Stannous hydrate dissolves in an excess of sodic hydrate, but it is insoluble in an excess of sodic carbonate and of ammonic hydrate and carbonate.

MERCURIC CHLORIDE changes stannous chloride, SnCl_2 , into STANNIC CHLORIDE, SnCl_4 , and a *white* precipitate of MERCUROUS CHLORIDE, Hg_2Cl_2 , is formed. *No other metal in solution gives this reaction with mercuric compounds.*

STANNIC COMPOUNDS.



SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate stannic compounds as STANNIC SULPHIDE, SnS_2 , *light yellow*. Stannic sulphide dissolves readily in ammoniac sulphide, and in sodic hydrate, and is precipitated completely from the solution, when a dilute acid is added gradually until the reaction becomes strongly acid. Stannic sulphide is nearly insoluble in ammoniac carbonate.

SODIC and AMMONIC HYDRATES and CARBONATES precipitate stannic compounds as STANNIC HYDRATE, $\text{Sn}(\text{HO})_4$, *white*. Easily soluble in sodic hydrate.

MERCURIC CHLORIDE does not give a precipitate with stannic chloride.

METALLIC ZINC precipitates TIN as crystalline metallic particles from stannic compounds in an acid solution, and the tin can be easily recognized by dissolving the metallic particles, after they have been washed by decantation. They are dissolved by heating them with a few drops of strong chlorhydric acid, and the stannous chloride thus obtained gives the precipitate, above described, with mercuric chloride.

ANTIMONY.



Metallic antimony is scarcely attacked by chlorhydric acid. It is oxydized when heated with moderately strong nitric acid to ANTIMONIC ACID, Sb_2O_5 , which is almost completely insoluble in nitric acid, but is readily soluble in hot concentrated chlorhydric acid. Aqua regia dissolves metallic antimony as ANTIMONIC CHLORIDE, SbCl_5 .

WATER precipitates solutions of antimony containing chlorhydric acid, particularly those of antimonious compounds, as a BASIC CHLORIDE. The precipitation can be prevented, or the precipitate dissolved by the addition of a sufficient quantity of acid; for this purpose tartaric acid is the most suitable.

METALLIC ZINC partially precipitates METALLIC ANTIMONY from its acid solutions, and when a piece of platinum in contact with a piece of zinc is introduced into solutions of antimony containing an excess of chlorhydric acid, metallic antimony is deposited upon the platinum as a *dark-brown stain*. *No other metal produces the same stain under like circumstances.*

ANTIMONIURETTED HYDROGEN. To obtain this body, follow exactly the directions given for obtaining arseniuretted hydrogen. (See page 56.)

The mirror obtained with antimoniuiretted hydrogen consists of a *black, sooty*, metallic deposit. It dissolves very slowly in hypochlorite of sodium. If the deposit is moistened with *yellow ammonic sulphide*, an *orange yellow* stain appears on the spot, when it is dried. These reactions are unimportant as tests for antimony, but a knowledge of them is necessary, in order that they may not be mistaken for evidences of the presence of arsenic.

BLOWPIPE REACTIONS. Antimony compounds, mixed with sodic carbonate and potassic cyanide, are quickly reduced, at a comparatively low temperature, to METALLIC ANTIMONY, *brittle shining grains*. The metal is completely volatilized by a strong heat. Metallic antimony, when heated on charcoal, burns with a white smoke, and gives an incrustation which is deposited at a great distance from the place heated, and is very volatile before the blowpipe flame.

OXYDATION BY SODIC NITRATE. When a sulphide of antimony is oxydized by a mixture of sodic carbonate and nitrate, ANTIMONIATE OF SODIUM, Na_3SbO_4 , insoluble in water, is formed.

ANTIMONIOUS COMPOUNDS.

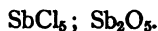
SbCl_3 ; $\text{KSbOC}_4\text{H}_4\text{O}_6$, *Tartar Emetic.*

SULPHYDRIC ACID and AMMONIC SULPHIDE

precipitate antimonious compounds, as ANTIMONIOUS SULPHIDE, Sb_2S_3 , *orange red*. The precipitation should be made in a cold solution containing tartaric acid and very little free chlorhydric acid. Antimonious sulphide dissolves in sodic hydrate and in ammoniac sulphide, most readily in *yellow* ammoniac sulphide, and is precipitated completely from these solutions, when a dilute acid is added gradually until the reaction becomes acid. It is nearly insoluble in ammoniac carbonate. *Yellow* ammoniac sulphide converts it into ANTIMONIC SULPHIDE, Sb_2S_5 . (See below.)

SODIC and AMMONIC HYDRATES and CARBONATES precipitate antimonious compounds, as ANTIMONIOUS OXIDE, Sb_2O_3 , *voluminous, white precipitate*. The precipitation only takes place after the lapse of a considerable time in solutions containing tartaric acid. Antimonious oxide is soluble in sodic hydrate.

ANTIMONIC COMPOUNDS.



SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate antimonious compounds in acid solution, as ANTIMONIOUS SULPHIDE, Sb_2S_3 , *orange red*. The precipitation does not take place immediately; but first an *orange* color appears; and it is only after passing sulphuric acid for a long time through the solution, that all the antimonious sulphide is precipitated. The precipitation should be made in a cold solution containing tartaric acid and very little free chlorhydric acid. Antimonious sulphide has the same properties as antimonious sulphide.

Antimonic oxide, Sb_2O_5 , plays the part of an acid with bases, and forms insoluble compounds with SODIUM, and soluble compounds with POTASSIUM.

Compounds of antimony can best be recognized by the stain, which they form on platinum, when the metal is precipitated by zinc from their solutions. (See page 54, Antimony.)

ARSENIC.

As.

Metallic arsenic is readily oxydized to arsenious or to arsenic compounds by nitric acid and is dissolved. It is not dissolved by chlorhydric acid. Metallic arsenic volatilizes at a heat below redness in a tube or on charcoal, and produces an odor like garlic, very characteristic of compounds of arsenic. The same odor is produced by heating a dry compound of arsenic on charcoal.

ARSENIURETTED HYDROGEN. To obtain this body, and to use Marsh's test, provide a 4 oz. flask with a tight-fitting cork, into which a funnel-tube and a small tube drawn off to a point and bent at a right angle,* are introduced. Fill it two-thirds full of dilute sulphuric acid, and add several pieces of pure zinc. A brisk evolution of hydrogen only commences after a few minutes. When that point is reached, wait five minutes for the expulsion of the air contained in the flask (without this precaution there is danger of an explosion), and light the hydrogen issuing from the point of the bent tube. The opening of the point and the quantity of gas evolved should be such, that the hydrogen burns with a blunt flame, about $\frac{1}{4}$ in. in length. Hold a bit of porcelain in the flame, in order to be certain that the reagents employed are free from arsenic. (See below.) If this is the case, pour a little of an acid liquid containing arsenic into the funnel-tube, while the hydrogen flame continues lighted. The flame after a few moments becomes white, and leaves a black stain of metallic arsenic upon a cold porcelain object held in it, in the same way that the flame of a candle would leave a deposit of soot on a cold surface. The arsenic stain or mirror is *shin-*

* In order to dry the gas, it is better to fit to the cork a chloride of calcium tube, bent downwards, so that it will not tip the flask over by its weight, and to adapt to the chloride of calcium tube, by means of a cork, a suitable jet directed upwards.

ing black (not *sooty black*, like antimony). It dissolves quickly in hypochlorite of sodium. When the stain is moistened with yellow ammoniac sulphide solution, and the spot is dried, it is *bright yellow*.

REACTION WITH THE AID OF HEAT. All dry compounds of arsenic, except some compounds with the higher metals, when they are mixed with dry sodic carbonate and potassic cyanide, and heated in a sealed tube, give a sublimate of metallic arsenic, which can be best recognized by breaking the sealed end of the tube after the formation of the sublimate, and by heating the sublimate quickly until it begins to volatilize, and by smelling of the upper end of the tube. A smell of garlic is proof of the presence of arsenic.

ARSENIOUS COMPOUNDS.



Arsenious acid is sparingly soluble in water. It dissolves more readily in sodic hydrate or carbonate, or in chlorhydric acid.

SULPHYDRIC ACID and AMMONIC SULPHIDE precipitate arsenious compounds from acid solutions as ARSENIOUS SULPHIDE, As_2S_3 , *yellow*. Arsenious sulphide dissolves very readily in sodic hydrate and ammoniac sulphide, and it also dissolves, although less readily, in ammoniac carbonate. It is precipitated from these solutions, when a dilute acid is slowly added until the reaction becomes acid.

SODIC and AMMONIC HYDRATES and CARBONATES produce no precipitate in arsenious compounds.

ARSENIC COMPOUNDS.



By heating any compound of arsenic with strong nitric acid, or by fusing a dry compound with a mixture of dry sodic carbonate and nitrate in a porcelain crucible, arsenic acid, As_2O_5 ,

or arseniate of soda, Na_2AsO_4 , is formed. These compounds are soluble in water.

SULPHYDRIC ACID and AMMONIC SULPHIDE do not immediately precipitate acid solutions of arsenic acid.

The solution (which should contain free chlorhydric acid,) at first becomes yellow, when a current of sulphydric acid is passed through it, and finally a *yellow* precipitate is formed; reduction to the state of an arsenious compound takes place slowly, and at the end of several days all the arsenic is precipitated as **ARSENIOUS SULPHIDE**, As_2S_3 , *yellow*.

See the properties of arsenious sulphide above.

ARGENTIC NITRATE produces no precipitate in acid solutions of arsenic compounds. (When chlorhydric acid is present a white precipitate of argentic chloride forms, and must usually, after the addition of an excess of argentic nitrate, be separated by decantation and filtration.) If sodic hydrate is added to a clear solution containing arsenic acid, As_2O_5 , and containing also an excess of argentic nitrate, until a permanent precipitate forms, and if then acetic acid is added until the reaction becomes acid, **ARGENTIC ARSENIATE**, Ag_3AsO_4 , *brick red*, is precipitated. This precipitate is soluble in ammonic hydrate and in dilute nitric acid.

This is the usual test for arsenic.

GOLD.

AuCl_3 .

Metallic gold is insoluble in any single acid, but it dissolves readily in aqua regia.

FERROUS SULPHATE solution in considerable quantity precipitates **METALLIC GOLD**, Au, *purplish-brown powder*, from acid solutions containing gold.

STANNOUS CHLORIDE solution precipitates **METALLIC GOLD**, Au, *purple flakes, very finely divided*, from acid solutions containing gold. (A very dilute solution of stannous chloride should be added, drop by drop, to the gold solution.)

SULPHYDRIC ACID gas slowly precipitates gold compounds in acid solutions as SULPHIDE OF GOLD, Au_2S_3 , *brown flakes*. The precipitate is soluble after a long digestion with *yellow* ammoniac sulphide.

BLOWPIPE REACTION. Gold is easily separated from its compounds with non-metallic elements by heating them on charcoal. It can be recognized as bright yellow globules.

Gold can easily be separated from the other metals and recognized by its precipitation with ferrous sulphate, or by its insolubility in any single acid.

Mix together solutions containing tin, antimony and arsenic, dilute the solution and add a considerable quantity of dilute chlorhydric acid. Pass sulphydric acid through the solution for an hour, and set it over night in a warm place. Filter the precipitate, wash and dry it completely. It contains all the metals in the form of sulphides; separate them according to Part III (79), and the following tests.

TESTS FOR ACIDS.

GROUP I.

ARSENIOUS, ARSENIC, CHROMIC, SULPHURIC, PHOSPHORIC, BORACIC, OXALIC, FLUORHYDRIC, CARBONIC, AND SILICIC ACIDS.

Acids which are precipitated from neutral or slightly alkaline solutions by baric chloride.

SECTION I.

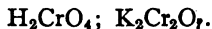
ARSENIOUS, ARSENIC AND CHROMIC ACIDS.

Acids which are precipitated as sulphides or reduced to an oxide by sulphydric acid.

ARSENIOUS and ARSENIC acids are precipitated as sulphides, and must always be detected by sulphydric acid.

Arsenic with this reagent plays the part of a metal. (See page 57.)

CHROMIC ACID.



Solutions containing chromic acid have a *yellow* color. Those which have an acid reaction are of a deeper color, and very concentrated acid solutions are *red*. Solutions containing chromic acid can usually be recognized by these colors. Chromic acid can easily be reduced to chromic oxide (see page 36), by boil-

ing its solution with chlorhydric acid and alcohol, or by adding to the solution sulphydric acid or ammoniac sulphide. Therefore, in all solutions to which these latter reagents have been added, chromic acid is changed to chromic oxide, and must be looked for among the metals.

BARIC CHLORIDE precipitates chromic acid in neutral solutions, or in solutions of which the only free acid is acetic acid (*i. e.* solutions which have been made alkaline by sodic hydrate, and acid by acetic acid), as **BARIC CHROMATE**, BaCrO_4 , *light yellow*. Chromium may be discovered in the precipitate of baric chromate thus obtained, by dissolving some of it in the borax bead. The *green color* of chromic oxide appears, even though very little chromate of barium was contained in the precipitate.

No other acid, which precipitates baric chloride, gives the same color in the borax bead.

SECTION II.

SULPHURIC AND SULPHUROUS ACIDS.

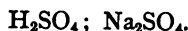
Acids which are precipitated by baric chloride in acid solutions, either immediately, or after oxydation.

BLOWPIPE REACTIONS. All compounds containing sulphur, when they are pulverized, moistened, and mixed with sodic carbonate, and when the mixture is heated on charcoal, form sodic sulphide. The sodic carbonate must be heated until it soaks into the charcoal, then if the portion of the charcoal, which has absorbed it, is dug out with a knife, moistened with water and laid on a piece of bright silver, the presence of sulphur may be detected by the appearance of a black stain of sulphide of silver. If no silver is at hand, the sodic sulphide can be extracted by soaking with water the portion of the charcoal which has been heated. The solution, after it has been filtered, and acidified with acetic acid, gives a precipitate of

SULPHIDE OF LEAD, PbS, *black*, with a solution of ACETATE OF LEAD.

For the separation of sulphuric and sulphurous acids, see Part III, (119) and (120).

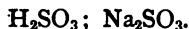
SULPHURIC ACID.



BARIC CHLORIDE precipitates compounds of sulphuric acid in neutral or acid solutions as BARIC SULPHATE, BaSO_4 , *white powder*. Baric sulphate is insoluble in chlorhydric, nitric, sulphuric, and the weaker acids. No other acid gives a like precipitate with BaCl_2 in acid solutions. COMPOUNDS OF LEAD AND STANNIC OXIDE *are the only others precipitated by sulphuric acid*.

CALCIC CHLORIDE in concentrated solutions gives a precipitate of CALCIC SULPHATE, CaSO_4 , but calcic sulphate is soluble in 380 parts of water.

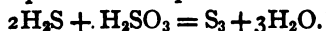
SULPHUROUS ACID.



Sulphurous acid, in acid solutions, can be recognized by its smell. Sulphurous acid is evolved from its acid solutions, when they are heated.

BARIC CHLORIDE precipitates compounds of sulphurous acid in neutral solutions as BARIC SULPHITE, BaSO_3 , *white powder*. Baric sulphite is decomposed and dissolved by chlorhydric and nitric acids. POTASSIC DICHROMATE oxydizes sulphurous acid in a solution, made acid with chlorhydric acid, to SULPHURIC ACID, $2(\text{H}_2\text{CrO}_4) + 3(\text{H}_2\text{SO}_3) = \text{Cr}_2(\text{SO}_4)_3 + 5\text{H}_2\text{O}$.

SULPHYDRIC ACID decomposes sulphurous acid in acid solutions with precipitation of sulphur.



THE IRON REDUCTION TEST. POTASSIC FERRICYANIDE AND FERRIC CHLORIDE, in consequence of the reduction of the ferric chloride, give a *blue* precipitate when these solutions are mixed, and a drop of the mixture is held on the end of a glass rod in an atmosphere containing sulphurous acid. Sulphydric acid gives the same reaction, and when sulphydric acid is present it is necessary to add to the solution sufficient plumbic acetate to precipitate it, before applying this test for sulphurous acid. This reaction is produced equally well in a solution; but there are many other reducing agents which act in the same manner, but which do not take the gaseous form, and for this reason it is best to add chlorhydric acid to the solution, to warm it, and to test the gas evolved for sulphurous acid with potassic ferricyanide and ferric chloride.

Sulphurous acid is usually recognized by its smell, or by the above test.

SECTION III.

PHOSPHORIC, BORACIC, OXALIC, AND FLUORHYDRIC ACIDS.

Acids which are precipitated by baric chloride in neutral solutions, but which are not precipitated in acid solutions.

For the tests by which these acids can be most easily detected see Part III, (121-4).

PHOSPHORIC ACID.



ALL THE METALS OF GROUPS II, III, IV, V precipitate compounds of phosphoric acid in neutral or in slightly alkaline solutions as PHOSPHATES. The precipitate is soluble in acids. The phosphoric acid cannot usually be separated from the metal by treating the precipitate with ammoniac hydrate, and from most metals it can only be partially separated by a treatment with sodic hydrate or carbonate. **MAGNESIC SULPHATE**

solution, to which a solution of ammoniac chloride, and of ammoniac hydrate is added, precipitates compounds of phosphoric acid as MAGNESIAC AMMONIAC PHOSPHATE, MgNH_4PO_4 , *white crystalline powder*, soluble in acids but insoluble in ammoniac hydrate. When the quantity of phosphoric acid is very small, the precipitate only forms after some time. *This is the best reagent for phosphoric acid in neutral or slightly alkaline solutions.*

MOLYBDATE OF AMMONIUM* precipitates compounds of phosphoric acid in an acid solution as a *yellow crystalline phospho-molybdate of ammonium*. No other acid gives a like precipitate. *The best reagent for phosphoric acid in acid solutions is molybdate of ammonium, when there are bases present which are precipitated by ammoniac hydrate.*

Sulphydic acid and ferrocyanhydric acid also precipitate ammoniac molybdate solution. (See part III, 123.)

BORACIC ACID.



ALL THE METALS OF GROUPS II, III, IV, V precipitate compounds of boracic acid in neutral or slightly alkaline solutions as BORATES. The precipitate is soluble in acids.

The precipitation of boracic compounds by metals of these groups is mostly incomplete, and the acid can be separated from the metal in almost all cases by treating the precipitate with ammonia or with sodic hydrate, or carbonate.

FLAME TEST. Boracic acid in dry compounds or in concentrated solutions imparts a beautiful *green color* to the flame of burning alcohol.

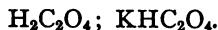
In order to perform the test, mix the dry boracic acid com-

* It is best to pour a few drops of the solution to be tested into the ammoniac molybdate solution instead of adding the ammoniac molybdate to the solution to be tested.

pound with a few drops of strong sulphuric acid in a small evaporating dish, add alcohol, and set the alcohol on fire. Stir the contents of the dish constantly during the combustion of the alcohol, and observe the *green color*, when the alcohol has mostly burned away.

TURMERIC PAPER TEST.* Boracic acid compounds, when their solution has been rendered acid by chlorhydric acid, turn a piece of turmeric paper, which has been dipped in the solution and completely dried, *brownish red*. *This is the usual test for boracic acid.*

OXALIC ACID.



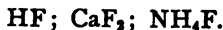
ALL THE METALS OF GROUPS II, III, IV, V precipitate salts of oxalic acid in neutral solutions as OXALATES except chromic oxide compounds. The precipitate is soluble in acids and in the case of many metals the oxalic acid is removed from the precipitate by ammoniac hydrate and by sodic hydrate and carbonate.

SULPHATE OF CALCIUM precipitates oxalic acid and its salts in a solution to which sufficient ammoniac hydrate to render the solution strongly alkaline, and then sufficient acetic acid to render the reaction acid, have been added, as CALCIC OXALATE, CaC_2O_4 , *white powder*. *No other acid, except fluorhydric acid, produces a precipitate under these circumstances, and calcic fluoride cannot easily be mistaken for calcic oxalate. (See fluorhydric acid, below.)*

SULPHURIC ACID (concentrated) decomposes dry compounds, or highly concentrated solutions of compounds of oxalic acid, with evolution of CARBONIC OXIDE and CARBONIC ACID. *Effervescence takes place.* $\text{H}_2\text{C}_2\text{O}_4 = \text{CO} + \text{CO}_2 + \text{H}_2\text{O}$.

* If iron is present it is necessary to boil the solution with sodic carbonate in excess, to filter and to use the filtrate for the turmeric paper test.

FLUORHYDRIC ACID.



Fluorhydric acid cannot be present in acid solutions in glass vessels. BARIUM AND CALCIUM SALTS in neutral solutions precipitate fluorhydric acid as BARIC OR CALCIC FLUORIDE, BaF_2 or CaF_2 . Baric fluoride is a voluminous white precipitate. Calcic fluoride is a gelatinous transparent precipitate, whose formation it is very difficult to observe.

It is usually unnecessary to test for fluorhydric acid except in solid substances.

SULPHURIC ACID (concentrated) sets fluorhydric acid free from its solid compounds, and the acid may be recognized by its property of etching glass. To perform the test, mix the pulverized substance, containing fluorine, with strong sulphuric acid in a lead cup,* or in a platinum crucible. Prepare a piece of glass by melting wax on it and pouring off all that does not adhere, leaving a thin coating of wax on the surface; scratch lines in the wax, laying the surface of the glass bare; cover the vessel, containing the fluorine compound and sulphuric acid, with the glass, and warm gently. After 15 minutes warm the glass, and rub off the wax; the surface exposed by the scratches will be etched by the fluorhydric acid.

Fluorhydric acid must always be removed, if present, by heating with strong sulphuric acid, before the other acids of Group I, except sulphuric acid and carbonic acid, are tested for in a substance.

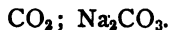
* A lead cup may be made by hammering up the end of a piece of 1/2 inch lead pipe until it is entirely closed, and by sawing off the pipe 1 1/4 inches from the end. Such a cup may be warmed on a sand-bath sufficiently for making the test without danger of melting it.

SECTION IV.

CARBONIC ACID AND SILICIC ACID.

Acids which are precipitated by baric chloride in neutral solution, but which are set free by acids, and which cannot be present in a solution that has been evaporated with an excess of an acid.

CARBONIC ACID.



Carbonic acid can only be present in considerable quantity in a solution which has an *alkaline reaction*.

ALL THE METALS OF GROUPS II, III, IV, V precipitate alkaline carbonates.

EVERY ACID sets free CARBONIC ACID, CO_2 , from its solutions. An effervescence or the formation of bubbles can be observed, when an acid is added to a solution of a salt of carbonic acid.

Carbonic acid can easily be recognized by the white precipitate of CARBONATE OF CALCIUM, CaCO_3 , which the gas produces in a drop of lime-water,* held in it on the end of a glass rod.

SILICIC ACID.



Silicic acid combined with bases in a solution is set free by all acids, and often separates as a precipitate; frequently, however, it remains for months in a solution after an acid has been added.

*The lime-water is soon destroyed by the absorption of carbonic acid from the air, and before using it to test for carbonic acid, the student should assure himself, that a drop of it gives a precipitate with the gas evolved from sodic carbonate, to which chlorhydric acid has been added.

When silicic acid has been separated from its combination with a base by the addition of an acid, and the solution has been evaporated completely to dryness, the silicic acid remains perfectly insoluble, when the dry mass is treated with water or acids to dissolve the bases. *This is the characteristic test for silicic acid.*

When silicic acid is present in a solution it is recognized and separated by the above process before any other tests are performed.

Many solid compounds of silicic acid are not acted upon by acids, and can only be brought into solution by the process described, Part III, XXIV.

The following properties of silicic acid must be considered, in order to determine whether a body, which has been left insoluble after a treatment with an acid, is silicic acid, or some other compound, which is likewise insoluble:—

Silicic acid is precipitated from its solutions by acids in the *gelatinous form*, or in the form of *amorphous white flakes*. Silicic acid, which has been dried, always has the latter form.

Silicic acid, after it has been fused with four parts of a mixture of sodic carbonate and potassic carbonate, forms a glass, which is entirely soluble in water. When the solution thus obtained is evaporated with an excess of nitric acid, and the dry mass is treated with water, no metal should go into solution, which gives a precipitate with sodic carbonate.

GROUP II.

CHLORHYDRIC, CYANHYDRIC, FERROCYNHYDRIC, FERRICYANHYDRIC, AND SULPHYDRIC ACIDS.

Acids which are not precipitated by baric chloride, but which are precipitated by argentic nitrate in nitric acid solution.

SECTION I.

CHLORHYDRIC AND CYANHYDRIC ACIDS.

Acids which give with argentic nitrate a white, flocculent precipitate, insoluble in dilute nitric acid, and are not precipitated by salts of iron in acid solution.

CHLORHYDRIC ACID.

HCl; NaCl.

PLUMBIC, MERCUROUS and ARGENTIC SALTS are the only compounds which give, with chlorhydric acid, precipitates insoluble in nitric acid.

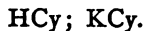
SULPHURIC ACID (concentrated) sets free chlorhydric acid from its compounds. Chlorhydric acid gas precipitates a drop of argentic nitrate, held on the end of a glass rod in an atmosphere containing it, as ARGENTIC CHLORIDE, AgCl, *white flakes*. Compounds containing cyanhydric, chloric and hypo-

chlorous acids, produce the same reaction. Chlorhydric acid gas does not bleach indigo solution. (See, however, Part III, 46).

ARGENTIC NITRATE precipitates solutions containing chlorhydric acid as ARGENTIC CHLORIDE, AgCl , *white flakes*, turning *purple* on exposure to light, *settling quickly after they have been shaken*. Argentic chloride is soluble in ammoniac hydrate, and completely insoluble in boiling nitric acid (concentrated). No other compound of silver except the ferro- and ferricyanide remains undissolved after this treatment. When ferro- or ferricyanhydric acid is present in a solution containing chlorhydric acid, follow the directions given, Part III, (130) (a) before applying the argenic nitrate test.

This is the usual test for chlorhydric acid.

CYANHYDRIC ACID.



The reactions of different classes of cyanhydric acid compounds must be considered separately.

SOLUBLE SIMPLE CYANIDES. Cyanides of metals of Groups I, II, AND III are soluble in water. Cyanhydric acid is set free from their solutions by even the feeblest acids (acetic and carbonic). (MERCURIC CYANIDE is soluble in water, but is not decomposed by alkalies nor by acids, except by sulphydric acid; and the tests described for cyanhydric acid cannot be applied to it. Sulphydric acid precipitates mercuric sulphide, and sets cyanhydric acid free.)

INSOLUBLE SIMPLE CYANIDES. Cyanides of metals of Groups IV and V, except mercuric cyanide, are insoluble in water, and the cyanides of metals of Group V are not decomposed, or are decomposed with great difficulty by acids. The insoluble cyanides dissolve readily in potassic cyanide, and the ordinary test for metals cannot be used with such solutions.

The cyanides can be precipitated from these solutions by the addition of an acid, with some exceptions, the two most remarkable of which are described separately. (See Ferro- and Ferricyanhydric Acids.)

FREE CYANHYDRIC ACID can be recognized by its smell, which is like that of bitter almonds. (The acid is very poisonous, and the fumes arising from a solution containing a considerable quantity of it should be inhaled with caution.)

ARGENTIC NITRATE precipitates soluble compounds of cyanhydric acid in an acid solution as ARGENTIC CYANIDE, *AgCy*, *white flakes, which do not settle so readily, when shaken, as argentic chloride, and which do not turn purple quickly in the light.* Argentic cyanide is wholly decomposed and dissolved by boiling a few minutes with strong nitric acid. It is also decomposed and cyanhydric acid goes into solution, when it is digested with dilute chlorhydric acid in contact with metallic zinc.

PRUSSIAN BLUE TEST. When a feebly acid solution containing cyanhydric acid is mixed with several drops of ferrous sulphate solution and with a drop of ferric chloride solution, and sodic hydrate is added until a precipitate forms, and the mixture is warmed for a minute, and then acidified with dilute chlorhydric acid, a *blue precipitate*, or more frequently a *blue coloration* appears, either immediately or after the addition of a drop of ferric chloride.

When ferro- or ferricyanhydric acid or both acids are present (see the following section), they must be removed from the solution, before the prussian blue test can be applied. To this end add to a small quantity of the solution an equal bulk of dilute sulphuric acid, and dilute with a considerable quantity of water; add ferric chloride or ferrous sulphate, or both together, according as ferro- or ferricyanhydric acid or both acids are present, and then add baric chloride until the blue precipitate appears of a much lighter shade; shake thoroughly, and allow the precipitate to settle for a few minutes, and filter. If only the first few drops run through the filter blue, they should be thrown away

and the remainder of the filtrate taken. If no clear filtrate can be obtained, add to the filtrate a little baric chloride and filter again. The filtrate is to be tested as above by the addition of sodic hydrate and, afterwards, of an acid for cyanhydric acid. A sufficiently capacious flask must be chosen for the operation. The only object in adding baric chloride is to facilitate the filtration from the blue precipitate.

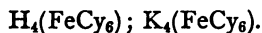
Cyanhydric acid is the only acid which gives this reaction under these circumstances.

SECTION II.

FERROCYANHYDRIC AND FERRICYANHYDRIC ACIDS.

Acids which give with argentic nitrate colored precipitates, which are not wholly destroyed on boiling with strong nitric acid, and which are precipitated by ferrous or ferric salts, and by cupric salts in dilute acid solutions.

FERROCYANHYDRIC ACID.



FERROUS SULPHATE precipitates ferrocyanhydric acid compounds in acid solutions as POTASSIC FERROUS FERROCYANIDE, $\text{K}_2\text{Fe}(\text{FeCy}_6)$, *bluish white* precipitate, which quickly turns *dark blue* through oxydation by the air.

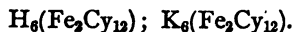
FERRIC CHLORIDE precipitates ferrocyanhydric acid compounds in acid solution as PRUSSIAN BLUE, $\text{Fe}_4(\text{FeCy}_6)_3$, *deep blue*.

Ferrocyanhydric acid is the only acid which gives this reaction.

CUPRIC SULPHATE precipitates ferrocyanhydric acid compounds in acid solution as CUPRIC FERROCYANIDE, $\text{Cu}_2(\text{FeCy}_6)$ *brownish-red powder*.

The metals are left as oxides, and the ferrocyanogen is dissolved as sodic ferrocyanide, when these precipitates are digested with sodic hydrate.

FERRICYANHYDRIC ACID.



FERROUS SULPHATE precipitates compounds of ferricyanhydric acid, in acid solution, as TURNBULL'S BLUE, $\text{Fe}_3(\text{Fe}_2\text{Cy}_{12})$, deep blue.

Ferricyanhydric acid is the only acid which gives this reaction.

FERRIC CHLORIDE does not precipitate compounds of ferricyanhydric acid in acid solution. The color of the solution is deepened.

CUPRIC SULPHATE precipitates compounds of ferricyanhydric acid in acid solution as CUPRIC FERRICYANIDE, yellowish-green powder.

The metals are left as oxides, and the cyanogen is dissolved as sodic ferricyanide, when these precipitates are treated with sodic hydrate.

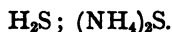
SECTION III.

SULPHYDRIC ACID.

An acid which gives a black precipitate with salts of lead, silver, copper, and many others in an acid solution.

No other acid gives a precipitate of the same color with these metals.

SULPHYDRIC ACID.



SULPHYDRIC ACID is set free from its solutions by all other acids except carbonic and cyanhydric acid, and it can be recognized by its smell. The sulphydric acid gas is given off with effervescence, when the solution is concentrated.

THE METALS OF GROUPS I and II form with sulphydric acid soluble sulphides, which have an alkaline reaction.

THE METALS OF GROUP IV, when the acid with which they are combined is neutralized, form with sulphydric acid **INSOLUBLE SULPHIDES**, which, with the exception of the sulphides of cobalt and nickel, are dissolved by cold dilute chlorhydric acid, with evolution of sulphydric acid.

METALS OF GROUPS V AND VI form with sulphydric acid insoluble sulphides, which are not decomposed by dilute acids. (See also Mercury, page 45 and page 47.)

LEAD PAPER TEST. A piece of paper moistened with plumbic acetate, and held over a solution from which sulphydric acid is set free by the addition of a stronger acid, is *blackened*. *No other acid gives this reaction.*

GROUP III.

NITRIC, CHLORIC, AND ACETIC ACIDS.

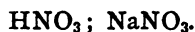
Acids which are not precipitated by any metal.

SECTION I.

NITRIC AND CHLORIC ACIDS.

Acids which deflagrate, when tested with the blowpipe on charcoal.

NITRIC ACID.



NITRIC ACID, when concentrated, is readily decomposed, when heated with copper turnings, and *red* fumes of NITRIC PEROXIDE, NO_2 , are given off. The reaction can be obtained with a moderately dilute solution, by adding to it concentrated sulphuric acid. No reaction is obtained with very dilute solutions.

FERROUS SULPHATE TEST. Add a few drops of a solution containing nitric acid to concentrated sulphuric acid in a test tube, and pour upon this solution a layer of cold ferrous sulphate solution. A *brown* or *red* color appears at the line of separation of the two solutions, arising from the absorption of nitrous gases by the ferrous sulphate.

This is the characteristic test for nitric acid.

CHLORIC ACID.



SULPHURIC ACID (concentrated). When a small quantity of a solid chlorate, or a very concentrated solution contain-

ing a chloric acid compound, is added to strong sulphuric acid, and heat is applied, a *peculiar yellow gas* (oxides of chlorine) is evolved, which has a characteristic suffocating odor, which precipitates ARGENTIC CHLORIDE in a drop of an argentic nitrate solution, and which bleaches a drop of an indigo solution, when these reagents are held on the end of a glass rod in an atmosphere containing the gas.

This is the characteristic test for chloric acid.

HYPOCHLOROUS ACID gives the same reactions as chloric acid, but that acid is easily set free, and evolved from its solution by dilute sulphuric acid, while chloric acid is not, and moreover it is usually present only in alkaline solutions.

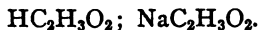
CHLORHYDRIC ACID in the presence of an oxydizing agent gives a similar reaction, but the yellow gas evolved (chlorine) is much less intense in color, and has a different odor. It is, however, very difficult to distinguish between the reaction given by chlorine in such a case, and that given by chloric acid compounds.

SECTION II.

ACETIC ACID.

An acid which does not deflagrate on charcoal.

ACETIC ACID.



THE STRONG MINERAL ACIDS set acetic acid free from its combinations.

Acetic acid can be recognized by the *odor of vinegar* peculiar to it.

SULPHURIC ACID TEST. When an equal bulk of alcohol is added to strong sulphuric acid, and a small quantity of a solution containing a compound of acetic acid is added, and the mixture is heated, a characteristic odor of acetic ether is given off.

In case gases are given off, which make it difficult to recognize the odor of acetic ether, it is advisable to provide the test tube in which the reaction is performed with a tube for distillation,* and to distil a small quantity of the alcohol into another test tube, to mix the distillate with water, to neutralize it with sodic carbonate, and to warm it; the odor of acetic ether can then be recognized in the liquid which was distilled.

This is the characteristic test for acetic acid.

ARGENTIC and MERCUROUS NITRATES precipitate concentrated neutral solutions of acetic acid compounds as ARGENTIC AND MERCUROUS ACETATES, $\text{AgC}_2\text{H}_3\text{O}_2$ and $\text{Hg}_2(\text{C}_2\text{H}_3\text{O}_2)_2$, *white crystalline scales*. The precipitates are soluble in dilute nitric acid, and also in a large quantity of water.

* Bend a 3-16 inch tube of about 1 foot in length at an angle of about 80° , so that one arm shall only be $1\frac{1}{2}$ inches long. Fit a cork to the test tube, and insert the bent tube in a hole bored through the cork with a round file.

PART III.

PRELIMINARY TESTS WITH NON-METALLIC SOLIDS.

EXAMINATION IN A CLOSED TUBE.

Use a piece of hard glass tubing $\frac{3}{8}$ inch in diameter, closed at one end [see page 25 (37)] for this examination. Introduce the substance, pulverized, or in small pieces, into the tube, wipe the inside of the tube if necessary, with a bit of rolled filter-paper and heat the substance, gently at first, but eventually to the highest temperature attainable with the flame of a Bunsen's lamp, or with the blowpipe flame. Observe carefully the changes which occur.

NO CHANGE. The substance contains no organic matter,

- (1) no readily fusible body, no readily volatile body, and no water.

Pass to the Examination on Charcoal, (page 81).

WATER. Substances containing water (usually water of

- (2) crystallization) deposit a *film of moisture* in the upper part of the tube, when they are heated. If the water colors turmeric paper *brown*, AMMONIA is present.

ORGANIC MATTER. Substances containing organic mat-

- (3) ter, *blacken* and give off gases when they are heated.

Should the substance contain organic matter it must be burnt until the organic matter is completely destroyed *

* In some special cases, as in examinations for mercury and arsenic, other processes of analysis must be employed, for which larger works must be consulted.

by heating with the lamp or blowpipe, on platinum foil or on a bit of porcelain, or in a porcelain dish or crucible, before further analysis, commencing with the examination on charcoal (page 81), is proceeded with.

A GAS IS GIVEN OFF.

OXYGEN. May be recognized by its property of rekindling
(4) a glimmering match, held in the tube.

HYPEROXIDES, NITRATES and CHLORATES evolve oxygen.

Nitrates and chlorates also deflagrate on charcoal. See (14.)

SULPHUROUS ACID, SO_2 , can be recognized by its
(5) smell.

Some SULPHATES of higher metals, and many SULPHITES, evolve sulphurous acid, when they are heated.†

SULPHYDRIC ACID, H_2S , can be recognized by its smell
(6) and by its property of blackening lead paper. See (44).

Some alkaline SULPHIDES, containing water, evolve sulphydric acid, when they are heated.

CARBONIC ACID, CO_2 , can be recognized by its property of
(7) extinguishing a lighted or glimmering match held in the tube. See also (40).

Some CARBONATES lose carbonic acid, when they are heated.

HYPONITRIC ACID NO_2 , appears as *red* fumes.

(8) NITRATES of the higher metals evolve hyponitric acid when they are heated.

AMMONIA, NH_3 , can be recognized by its smell, and by its

† Many sulphides of higher metals give off sulphur in the form of sulphurous acid, when they are roasted with access of air. The sulphides, finely pulverized, may be heated red-hot in a tube, open at both ends, and held in an inclined position to favor the draught, and the sulphurous acid may be detected by its smell at the upper end of the tube.

- (9) property of turning moist turmeric paper *brown*. Salts of ammonia, in the presence of alkalies, and some organic substances, evolve ammonia, when they are heated.

[Better tests for these bodies, with the exception of oxygen, are given in the following pages, since it is often difficult to observe the formation of a gas in a small tube; the phenomena described above should, however, be looked for when substances are heated in a closed tube.]

A SUBLIMATE FORMS.

An opinion may be formed of the volatility of the sublimate, according to the distance from the heated part of the tube, at which it is deposited.

SULPHUR sublimes easily and solidifies in *reddish-brown* (10) drops, which become *yellow* or *yellowish-brown* on cooling.

Some METALLIC SULPHIDES give off a portion of their sulphur, when they are heated.

AMMONIC SALTS form white sublimate. Touch the sublimate with a drop of sodic hydrate, or with a bit of paper moistened with sodic hydrate, and if the smell of ammonia is given off, it consists of an ammoniac salt.

MERCURY. Metallic mercury sublimes as a *grey film*, (12) which augments to form *globules* when the quantity of mercury is large.

MERCURIC SULPHIDE, HgS , gives a *black* sublimate, which becomes *red*, when it is rubbed.

MERCUROUS CHLORIDE, Hg_2Cl_2 , and MERCURIC CHLORIDE, HgCl_2 , give a *white* sublimate, which turns *black*, when it is moistened with ammoniac sulphide solution.

ARSENIC. Metallic arsenic sublimes and deposits itself (13) as a *brilliant black metallic ring* in the tube.

ARSENIOUS OXIDE, As_2O_3 , forms a *white crystalline* sublimate, which turns *yellow*, when it is moistened with sulphydric acid solution.

ARSENIOUS SULPHIDE, As_2S_3 , forms a sublimate, which is *reddish yellow* when hot, and *yellow* when cold. It is somewhat less volatile than sulphur.

RECAPITULATION (10) to (13) SUBLIMATES.

The substance is heated in a closed tube.

WHITE SUBLIMATE—ammonic salts (11); mercurous chloride, Hg_2Cl_2 and mercuric chloride, HgCl_2 (12); and arsenious oxide, As_2O_3 (13).

YELLOW SUBLIMATE—sulphur (10); and arsenious sulphide, As_2S_3 (13).

BROWN SUBLIMATE (while hot)—sulphur (10).

REDDISH-YELLOW SUBLIMATE (while hot)—arsenious sulphide, As_2S_3 (13).

GRAY METALLIC SUBLIMATE—mercury (12).

BLACK SUBLIMATE—arsenic (13); and mercuric sulphide, HgS , red when rubbed (12).

EXAMINATION ON CHARCOAL.

Hollow out a small cavity in a piece of charcoal (see page 24) (34), and heat a portion of the solid substance with the blow-pipe flame.

NITRATES and CHLORATES enter into a vivid combustion, called *deflagration*, when they are heated on charcoal.

POTASSIUM and SODIUM SALTS melt, and some of (15) them are imbibed by the pores of the charcoal when they are heated.

COMPOUNDS OF THE METALS OF GROUP II and III. ALSO ZINC COMPOUNDS and SILICIC

- (16) **ACID**, remain as a *white infusible* mass on the charcoal, after heating. Frequently, when heat is first applied, they melt in their water of crystallization, and afterwards become solid.

ALUMINIC OXIDE becomes *blue*, and ZINC OXIDE becomes *green*, when they are moistened with cobaltic nitrate, and heated in the oxydizing flame.

SALTS OF THE METALS OF GROUPS IV and V

- (17) leave a dark-colored residue, when they are heated on charcoal. The oxides of these metals generally assume a darker color, when they are heated. Exceptions: Zinc and Mercury.

SALTS OF AMMONIA and MERCURY, also COM-

- (18) **POUNDS OF ARSENIC and ANTIMONY**, which do not contain another metal, volatilize completely, when they are heated on charcoal.

GOLD and SILVER COMPOUNDS, also OXIDES OF

- (19) **LEAD and BISMUTH**, give *bright metallic* globules, when they are heated on charcoal.

FUSION WITH SODIC CARBONATE.

- (20) When metals of groups IV, V, and VI appear to be present, (see 17) mix a small quantity of the pulverized substance with two or three times its bulk of sodic carbonate in the palm of the hand, moisten with water, and form the mixture, by working it with a knife-blade, into a ball the size of a pea. Place the ball in a cavity scooped out of a piece of charcoal, and heat with the inner blow-pipe flame, until almost all of the carbonate of soda has been imbibed by the charcoal. Many metals are reduced and appear as metallic globules in the cavity of the charcoal, and those which are volatile deposit an incrustation of their oxides on the charcoal. This incrustation is to be looked for at a greater or less distance from the cavity, according to the volatility of the metal, and always in the

direction in which the metallic vapors are blown by the flame.

The physical and chemical properties of the globules and the color of the incrustations afford means of recognizing several metals, usually, however, only when they are not associated with others.

IRON, COBALT, NICKEL, and MANGANESE COM-

(21) POUNDS give neither globule nor incrustation.

GOLD, SILVER, and COPPER COMPOUNDS give

(22) *malleable globules* which can be distinguished by the respective colors of the metals. They give no incrustation.

ZINC COMPOUNDS give no globules, but a *white incrusta-*

(23) *tion*, ZnO , near the spot heated. The incrustation is *yellow* while hot. It is not volatile in the oxydizing flame. It becomes *green* when it is moistened with nitrate of cobalt, and heated in the oxydizing flame.

TIN COMPOUNDS give very *ductile white globules*.

(24) The incrustation, SnO_2 , produced by tin compounds, is *dirty yellow* when hot, and *lighter* when cold. It is deposited in the immediate vicinity of the cavity, and it is very difficult to distinguish it from the ash of the charcoal.

LEAD COMPOUNDS give very *ductile globules*.

(25) The incrustation, PbO , is *bright yellow* when hot, and *pale yellow* when cold. It is deposited at a greater distance than SnO_2 from the cavity. When the blowpipe flame is directed upon the incrustation of PbO , it vanishes and the flame is colored *blue*.

BISMUTH COMPOUNDS give *brittle globules*. The in-

(26) *crustation*, Bi_2O_3 , is *orange yellow* when hot, and *bright yellow* when cold. It vanishes, when the blowpipe flame is directed upon it, but it does not impart a *blue* color to the flame.

ANTIMONY COMPOUNDS give *brittle globules*, but metallic (27) antimony is so volatile, that frequently these are driven off by the heat required for their reduction. Sometimes fumes, arising from the vapor of antimony, are visible. The incrustation Sb_2O_3 , is *white*. It is deposited at a greater distance than PbO from the cavity, and it can easily be driven from one place to another on the charcoal by the heat of the blowpipe flame.

ARSENIC COMPOUNDS give no globules, but a characteristic (28) *garlic odor*. The incrustation, As_2O_3 , is *white*, and it is still more volatile than Sb_2O_3 .

When the compound contains several metals, that can be reduced, they alloy with each other, and it is usually impossible to recognize the metals in the presence of each other by their physical properties; also, the incrustation given by one metal frequently obscures that given by another.

If a sufficient quantity of the metal can be easily reduced, it is always advisable to treat it with solvents in the manner to be described under metals. (See page 96.)

SULPHUR. The following modification of the fusion with (29) carbonate of sodium on charcoal is a valuable test to discover sulphur in baric sulphate and in sulphides. If a piece of the charcoal, which has imbibed the soda, is moistened and laid on a silver coin, a *black* stain appears, if the coin is washed after a few minutes, in case sulphur is present. The charcoal may also be pulverized and treated with water, and if the solution, after being filtered, gives a black precipitate with plumbic acetate solution, sulphur is present.

RECAPITULATION OF THE EXAMINATION ON CHARCOAL.

The substance is heated on charcoal.

DEFLAGRATION.—Nitrates and chlorates (14).

FUSION.—Potassium and sodium salts (15).

WHITE INFUSIBLE RESIDUE.—Compounds of metals of groups II and III; zinc salts and silicic acid (16).

DARK COLORED RESIDUE.—Compounds of metals of groups IV and V except zinc and mercury (17).

COMPLETE VOLATILIZATION.—Ammonic and mercuric compounds, and compounds of arsenic and antimony, which contain no other metal (18).

BRIGHT METALLIC GLOBULES.—Silver and gold compounds, and the oxides of lead and bismuth (19).

The substance is mixed with sodic carbonate and heated on charcoal.

METALLIC GLOBULES WITHOUT INCRUSTATION.—Gold, silver, and copper (22); tin (24).

METALLIC GLOBULES AND INCRUSTATION.—Lead (25); bismuth (26); antimony (27).

INCRUSTATION WITHOUT GLOBULES.—Zinc (23). GARLIC ODOR. Arsenic (28).

FORMATION OF SODIC SULPHIDE.—All compounds containing sulphur (29).

PRELIMINARY TESTS WITH METALLIC BODIES.

EXAMINATION IN A CLOSED TUBE. (See page 25 (37.)

MERCURY. Amalgams, containing mercury, give a sublimation (30) of METALLIC MERCURY, when they are heated. At first a *gray film* forms in the upper part of the tube, and when the amount of mercury is considerable, *fine globules* of metallic mercury are formed, which agglomerate and become more distinctly visible, when they are rubbed with a copper wire.

ARSENIC. Some metallic compounds, containing ARSENIC, (31) give a metallic *mirror* or ring in the upper part of the tube, when they are heated.

EXAMINATION ON CHARCOAL.

Heat a piece, half as large as a pea, of the metallic (32) substance in a cavity, on a piece of charcoal. See Part I, page 24 (34). The phenomena to be observed are the formation of the incrustations described (page 85 and 86), the smell of ARSENIC, and the vapors of MERCURY and ANTIMONY.

COPPER colors the blowpipe flame *green*, or in the presence of chlorine, *blue*.

PRELIMINARY TESTS WITH NON-METALLIC

SOLIDS (*continued*).

TESTS WITH THE BORAX BEAD.

If the substance to be tested appears to be the oxide, or an oxygen-salt of a higher metal (see 17), dissolve some of it in the borax bead.

COBALT colors the bead *blue* in the oxydizing and in the (33) reducing flame.

COPPER colors the bead *green* when hot, and *blue* when (34) cold, in the oxydizing flame. It colors the bead *red*, when cold in the reducing flame.

CHROMIUM colors the bead *green* in both flames. (35)

IRON colors the bead *brownish-red* when hot, and *yellow* when (36) cold, in the oxydizing flame.

NICKEL colors the bead *violet* when hot, and *pale-brown* when (37) cold, in the oxydizing flame. The color disappears in a *good* reducing flame. (See page 42.)

MANGANESE colors the bead *amethyst* in the oxydizing (38) flame. The color disappears in a good reducing flame.

(39) THE OXYGEN-COMBINATIONS OF THE REMAINING METALS color the bead very slightly or not at all.

CONCENTRATED SULPHURIC ACID TEST.

CONCENTRATED SULPHURIC ACID, with the aid of heat, sets free other acids from most of their combinations with metals, and frequently in such a form, that they can be recognized by simple tests. This reaction is not, of course, a method of separation; one acid may obscure the test for another, and the possible cases are so complicated, that it would be useless to attempt to describe them all; therefore, if the result of the sulphuric acid test appears doubtful, it is best to reserve judgment of its value, until after the tests for acids in solution have been applied.

The reactions of acids or their salts, when a small quantity of a solid substance or of a very concentrated solution is added to a few cubic centimetres of strong sulphuric acid in a test tube, are described below. Heat should be applied, after the reaction, which takes place at the ordinary temperature, has been observed.

CARBONIC ACID. *Effervescence.* Carbonic acid gas renders turbid a drop of lime-water held on the end of a glass rod in the test-tube (see page 67, foot-note). Carbonic acid is also detected by the chlorhydric acid test, and in many cases that test is preferable to the one with sulphuric acid, since oxalic acid does not give the same reaction with chlorhydric acid. See (68).

OXALIC ACID. When a dry compound of oxalic acid is (41) added to strong sulphuric acid, and the mixture is heated, carbonic acid and carbonic oxide are evolved.

The sulphate of calcium test for oxalic acid (121); is more accurate than that with sulphuric acid.

CYANHYDRIC, FERRO- and FERRICYANHYDRIC ACIDS. Compounds of these acids evolve, (41a) when perfectly dry, carbonic oxide with effervescence, when they are heated with strong sulphuric acid. Usually, however, a faint odor of cyanhydric acid can be detected.

The special tests are more valuable. See (69), and (127), (128), and (129.)

FLUORHYDRIC ACID. When a solid substance or a (42) concentrated solution, containing fluorine, is heated with strong sulphuric acid, fluorhydric acid is evolved, which etches glass (see page 66); when silicic acid or a silicate is present, fumes of fluoride of silicon, which give a precipitate of silica in a drop of water held over them on the end of a glass rod, are evolved.

If fluorhydric acid is discovered, the substance which is to be used for page 98, III, and the following tests must be heated with sulphuric acid in a platinum vessel until the fluorhydric acid is driven off completely.

A separate portion can be used for testing for sulphuric acid. See (29).

SULPHYDRIC ACID. Compounds, containing this acid, (43) evolve it (often with effervescence), when strong sulphuric acid is added to them. Sulphydric acid may be recognized by its smell and by its property of blackening paper, dipped in plumbic acetate solutions. See (70) and (125).

SULPHUROUS ACID. Compounds, containing this acid, (44) evolve it with effervescence, when strong sulphuric acid is added to them. Sulphurous acid, when free from sulphydric acid, and from some others, can be recognized by its smell. (See 71).

CHLORIC and HYPOCHLOROUS ACIDS. (See (45) page 76.) Compounds containing these acids evolve a *yellow* gas on the addition of strong sulphuric acid, even when the mixture is not heated. The gas can best be recognized by its color, its odor and its strong bleaching action on a drop of indigo solution, held in the tube on the end of a glass rod. This gas precipitates nitrate of silver.

CHLORHYDRIC ACID. Compounds containing chlorhy-

- (46) dric acid evolve the gas HCl , frequently with effervescence, on the addition of sulphuric acid. The gas does not bleach indigo solution, but precipitates nitrate of silver, held on the end of a glass rod in the tube. Chlorhydric acid, in the presence of an oxydizing agent, evolves chlorine under the same circumstances. The gas produces the same reactions as are produced by the gas evolved by chloric and hypochlorous acids, but it can be distinguished from them by its smell and by its color, which is a less intense yellow. For special test, see (126).

NITRIC ACID. Compounds, containing nitric acid in con-

- (47) siderable quantity, produce reddish fumes, when heated with sulphuric acid in the presence of copper turnings or of any other reducing agent. The following test for nitric acid is more delicate: Mix a little of the powdered substance or solution with strong sulphuric acid, and pour cautiously upon the acid a solution of ferrous sulphate. A *brown* or *red* color at the line of separation of the two solutions indicates the presence of NITRIC ACID, HNO_3 . See (130) (a) for this test in the presence of ferro- or ferricyanhydric acid.

ACETIC ACID gives with sulphuric acid an odor of vinegar.

- (48) The following test is more delicate: Add an equal volume of alcohol to strong sulphuric acid and then add the solid substance or concentrated solution, supposed to contain ACETIC ACID, $\text{C}_2\text{H}_4\text{O}_2$. If this acid is present the odor of acetic ether will be perceptible on heating the mixture. It is well to add pure acetic acid at the same time to a mixture of sulphuric acid and alcohol, in order to compare the odor produced with that observed in the test. If other gases render the odor of acetic ether difficult to perceive, the precautions described on page 77 must be observed.

RECAPITULATION OF SULPHURIC ACID TEST

A colorless gas is given off.

THE GAS IS WITHOUT ODOR.—Carbonic acid (40); oxalic acid (41).

A PUNGENT SUFFOCATING ODOR.—Fluorhydric acid (42); chlorhydric acid (46).

AN ODOR OF BITTER ALMONDS.—Cyanhydric, Ferrocyanhydric and ferriecyanhydric acids (41) (a).

AN ODOR OF ROTTEN EGGS.—Sulphydric acid (43).

AN ODOR OF BURNING SULPHUR.—Sulphurous acid (44).

AN ODOR OF VINEGAR.—Acetic acid (48).

A colored gas is given off.

The gas has also a peculiar suffocating odor.

YELLOW GAS.—Chloric and hypochlorous acids (45).

FAINT YELLOW GAS.—Chlorine (46).

RED FUMES—. Nitric acid with copper (47).

times with hot water.

The powder may contain antimony, tin and lead (sulphate). Transfer it to an evaporator and boil it with about 10 c.c. of a concentrated solution of tartaric acid; if the ppt is entirely dissolved lead and tin are absent. Test the solution with H_2S ; an orange ppt. indicates antimony.

If the white powder did not dissolve entirely, filter and test the filtrate for antimony with H_2S . Heat the insoluble residue gently with 10-15 c.c. of ammonium tartrate, filter if any soluble residue remains, and test the filtrate for lead with H_2S (black ppt.). As this ppt. by H_2S may also contain tin it should be reserved for treatment with NH_4HS in case this metal not found in the step immediately following this.

If a residue was left insoluble in ammonium tartrate, test it on a filter, wash, and dissolve in conc. HCl with the aid of heat, and test the solution for tin as directed in § 4.

§ 5 b. No white powder was formed in making the solution
To the nitric acid solution obtained (or to the filtrate from the white powder obtained in § 5 a) add dilute HCl , and if a ppt. is formed test it for lead, silver and mercury (suboxide) as directed in 60 and 64, page 99 of Craft's Ind. Analysis.

Treat the filtrate from this ppt., or the original HNO_3 solution, if no ppt. was formed by HCl , as directed in § 3.

§ 5 c. Preparation of a solution for treatment with H_2S (or with NH_4HS) if the student has not advanced as far as the 5th & 6th groups.

In all cases, before proceeding to test the solution with either of these reagents it must be so prepared that it will be moderately acid with HCl , and will not contain any free nitric acid, or great excess of strong acid of any kind. If the substance was dissolved in water the solution is already in this condition after it has been taken through the HCl test; or if it was dissolved in dilute HNO_3 the same is true.

But, otherwise, the excess of HNO_3 used for the solution of the substance must be replaced by HCl . Add a quantity of the latter about half as great as that of the nitric acid that was used. Boil the liquid down almost to dryness in the hood and add from 25 to 50 c.c. of water to the residue.

If the substance was dissolved in conc. HCl the solution should be evaporated down in the same manner to sp. gr.

METHODS OF DISSOLVING NON-METALLIC BODIES.

Substances to be tested may be divided into three classes.

1st CLASS : Bodies soluble in water.

- (49) To ascertain whether a body is entirely soluble, take a few grains of the substance, which, when it dissolves with difficulty, must be in as finely divided a condition as possible, and digest them with a considerable quantity of water in a test tube. If the substance does not dissolve completely, boil it for a few minutes with water. If it still does not dissolve completely, filter and evaporate a few drops of the filtrate on platinum foil, in order to see whether anything has dissolved.

If the substance is partly soluble in water, treat a considerable quantity in the manner directed above, repeat the boiling with water once or twice, wash thoroughly with water, and filter, taking care that as little as possible of the solid substance goes upon the filter, and use the insoluble portion for the following test.

A separate analysis should usually be made of the part of a substance, which is soluble, and of that which is insoluble in water.

If the substance is wholly insoluble in water, it may be used immediately for the following test. It should usually be finely pulverized.

2d CLASS : Bodies insoluble in water, but soluble in chlor-

- (50) hydric acid or in nitric acid or in aqua regia. Digest a few grains, or a very small quantity of the finely powdered substance with dilute chlorhydric acid. If it does not entirely dissolve, boil it with the acid. If it still does not entirely dissolve, pour off the liquid, and boil the insoluble substance with strong chlorhydric acid. If the substance is not wholly soluble in chlorhydric acid, repeat

the trial with nitric acid in the same way, using a fresh portion of the substance. If the substance is still insoluble, use a mixture of both acids (aqua regia).

If sulphur, which may be recognized by its color and its low specific gravity, or silicic acid, which may be recognized by its peculiar gelatinous aspect, separate out on the addition of chlorhydric or of nitric acids, the substance must be considered as soluble, and after filtration (in case of silicic acid, see (60) the solution must be tested according to page 98 (III) or page 101 (VI) accordingly as nitric or chlorhydric acid has been the solvent.

If the substance is only partially soluble after treatment as above, the insoluble portion must be washed carefully and separated by filtration from the soluble portion, and it must be subjected to the tests for the 3rd Class. See XXII.

3d CLASS: Bodies insoluble in water and in chlorhydric, nitric (51) and nitro-chlorhydric acids (aqua regia).

The tests to be applied to bodies of this class follow those in the scheme for testing bodies in solution. See XXII.

METHODS OF DISSOLVING METALLIC BODIES.

Metals are divided into three classes.

1st CLASS: Metals which are not attacked by NITRIC ACID. (52)

If the metal does not appear to be entirely soluble in dilute nitric acid, even after boiling, use strong nitric acid.

GOLD is insoluble, also alloys containing a very large proportion of gold.

2d CLASS: Metals which are attacked by nitric acid, and (53) converted into oxides (*white powder*), which are insoluble in the acid.

TIN and ANTIMONY belong to Class II.

3d CLASS: Metals which dissolve entirely in nitric acid. All

(54) the remaining metals dissolve entirely in nitric acid. With metals of this class the solution should be effected by heating with nitric acid in an evaporating dish, until no more red fumes are given off. The greater part of the acid should then be evaporated, and water should be added, and the solution should be tested according to (page 98, III). If the metal remains partly insoluble, a small portion of it should be carefully tested by boiling with strong nitric acid, to see whether it will not dissolve by using a sufficient quantity of the acid.

(55) An alloy may contain metals of each class, therefore, after the treatment with nitric acid described in the last paragraph, if an insoluble residue is found it should be dissolved in aqua regia, and the solution boiled, until the smell of chlorine ceases to be given off. Then if the residue had a metallic appearance the solution must be tested for gold. See (84). If it was a white powder it must be tested for tin and antimony. See (82) and (83).

SOLUTION IN CHLORHYDRIC ACID.

(56) Sometimes a test shows that the metallic body can be readily dissolved in chlorhydric acid, (see ZINC, page 37 and IRON, page 39,) and in this case the chlorhydric acid solution should be preferred, and it should be tested according to page 103, VI.

TESTS FOR METALS AND ACIDS.

BODIES IN SOLUTION.

CLASSES I AND II. (See page 91).

Substances dissolved in water or in acids.

(The tests, I & II, which follow, are only to be used for substances dissolved in water, or where the solvent is unknown.)

I. REACTION WITH TEST PAPER.

(57) Observe the reaction with litmus paper.

(58) If the reaction is acid, add to a small portion of the solution sodic carbonate drop by drop, until the effervescence ceases, and then heat to boiling.

If metals of the 2d, 3d, 4th, and 5th groups are present, a precipitate will be formed, except in a few special cases.

Some idea of the amount of free acid present in the solution, can be formed by observing the amount of sodic carbonate required to neutralize it.

II. EVAPORATION.

(59) Test one or two drops of the solution by evaporation on platinum foil (or on a bit of glass or porcelain, if there is danger of injury to the platinum), to see whether it leaves a residue.

If this test shows the presence of solid matter in the

solution, a portion of the latter may be evaporated to dryness in a small porcelain dish, and some of the tests for solids, particularly the EXAMINATION ON CHARCOAL (page 81), and the CONCENTRATED SULPHURIC ACID TEST (page 87), may be applied to the dry substance thus obtained. It is best only to use the evaporation and tests applied to the dry substance, to settle any doubts that remain in regard to the constitution of the solution, after the tests usually applied to solutions have been performed. (See page 100, &c.)

When a solution of unknown origin is presented for analysis, it should always be heated in order to see whether a gas is given off, that can be recognized by the tests (page 87, and page 88).

SILICIC ACID is usually not present, except in alkaline solutions; it occurs, however, in small quantities in spring and river waters, and it may also exist in acid solutions.

Unless silicic acid is known to be absent it should be tested for, and removed before making the remaining tests. To this end render the solution acid with chlorhydric acid, and evaporate carefully to dryness in a small porcelain dish. Care must be taken not to heat the dish over the lamp, after its contents have become dry. Treat the dry substance with a few drops of acid, and then boil with water. If there is an insoluble residue, it consists of silica, SiO_2 . (See page 68.)

Subject the chlorhydric acid solution to the remaining tests, beginning with page 101, VI; and if there is reason to suspect that the precipitate, left after the treatment with an acid, contains other substances besides silica, examine it according to page 98, IV.

TESTS FOR METALS.

The following scheme of testing for metals is founded upon the *successive* precipitation of a number of groups, which include all the metals as far as Group II. After the metals of the higher groups have been removed by precipitation, or have been found to be absent, those of Group II are precipitated successively. Sodium and potassium are detected by the colors of their flames in a solution, which has been freed from all the higher metals, except magnesium. Ammonia can be detected in a solution without having reference to its other constituents.

After the separation into groups has taken place by precipitation with the general reagents, each precipitate, which may contain one or all the metals belonging to its group, is usually dissolved, and the further analysis is performed by testing in the several solutions for all the metals which they may contain. These special tests sometimes require the separation of the different metals, one after another, in a particular order, while sometimes a test for a metal may be applied to the solution without regard to the presence of other metals. In all cases the conditions requisite for applying the tests will be described.

The general tests must be applied in the following order: 1st, Chlorhydric acid to effect the precipitation of silver and mercurous compounds, and of lead if it is present in large quantity. 2nd, Sulphydric acid in an acid solution to precipitate small quantities of lead and the metals of Group V, Section II, and of Group VI. (From this precipitate the metals of Group VI are separated by dissolving them in ammoniac sulphide). 3rd, Ammoniac hydrate until the reaction becomes alkaline, ammoniac

chloride and ammoniac sulphide to precipitate the metals of Groups III and IV.

Whenever a single metal or a group of metals is precipitated, some of the liquid containing the precipitate must be poured on a filter, and the first drops of the solution, which run through, must be tested with some of the reagent, which was used to produce the precipitation, in order to ascertain whether it has been completely effected. Should a fresh precipitate make its appearance, every thing must be poured back from the filter into the test tube or flask and more of the reagent must be added. This operation must be repeated until no precipitation is produced by the same reagent in the liquid which runs through the filter. After a little practice it is easy to estimate how much of a reagent is required to effect a complete precipitation. A surplus over this quantity is called an excess of the reagent.

When it has been found that a slight excess of the reagent has been added, the whole of the liquid and the precipitate together must be poured upon the last filter and the liquid must be allowed to drain off. The liquid is to be tested for metals of the succeeding groups, and the precipitate must usually be completely freed from it; otherwise the separation has no value. To this end, water must be blown on the precipitate from the wash-bottle and allowed to drain off. Care must be taken not to let the water overflow the edge of the filter. The washing must be continued until the water, which flows through the filter, is proved, either by evaporation on platinum foil or by the application of tests for the succeeding groups, not to contain any metal in solution. The last portions of the wash-water, which pass through the filter, may be thrown away, as they contain very little of the substances to be tested for.

The value of analyses depends upon the care with which the separation of precipitates from the liquid in which they have been, formed, is executed. For special directions for filtering, see pages 22 and 23, (29) and (30).

When the reaction of a liquid in a test-tube is to be tested, always close the tube and shake it thoroughly before dipping the test paper in.

METALS OF GROUPS VI & V.

III.—CHLORHYDRIC ACID TEST.

Metals in acid solution.

GROUP V., SECTION I.

In case the solution is known to contain chlorhydric acid, pass to page 101 (VI).

If the solution has an alkaline reaction, pass to page 99 (IV).

- (61) Add to a very small quantity of the solution to be tested a few drops of dilute chlorhydric acid. If a precipitate forms, continue to add chlorhydric acid drop by drop, as long as it seems to increase in quantity, then add a quantity of chlorhydric acid about equal to that already added.

If no precipitate is formed, or if it is dissolved on further addition of dilute chlorhydric acid LEAD IN LARGE QUANTITY, SILVER, AND MERCUROUS SALTS ARE ABSENT. *Pass to page 101 (VI).*

- (62) If a permanent precipitate is formed, it may consist of PLUMBIC, ARGENTIC, AND MERCUROUS CHLORIDES. Most of the succeeding tests for the detection of bases must be performed in a solution freed from these metals. Therefore, if a precipitate has been observed, treat a considerable quantity of the solution, in the same way that the small portion was treated, and after an excess of chlorhydric acid has been added, shake the liquid for a minute or two. The precipitate will then settle in a short time, leaving the solution nearly clear. The solution should be

decanted through a filter, and the precipitate washed twice by decantation through the filter with water acidulated with chlorhydric acid.

Test the filtrate according to page 101 (VI).

LEAD. Add a small quantity of water to the precipitate, and (63) boil, then let the precipitate settle, and decant the clear liquid through the same filter, which was used in (62) into another vessel. Add to the filtrate an equal bulk of alcohol and a small quantity of dilute sulphuric acid. If a white precipitate forms, it consists of SULPHATE OF LEAD, $PbSO_4$.

If lead is found, the precipitate must be washed as before, with boiling water by decantation, until the filtrate gives no *black* precipitate with ammoniac sulphide. *If no precipitate remains after the washing, NO ARGENTIC OR MERCURI- OUS SALTS are present, Pass to page 101 (VI).*

SILVER. If a precipitate remains, after washing with boiling (64) water, add to it ammonia, pour the solution through a filter, and acidify with nitric acid. If a precipitate forms, it consists of ARGENTIC CHLORIDE, $AgCl$.

MERCURI- OUS SALTS. If a precipitate remains, after (65) ammonia has been added, it will have a *gray or black color*. The precipitate consists of A MERCURI- OUS COM- POUND OF AMMONIA.

IV.—CHLORHYDRIC ACID TEST.

Metals in alkaline solutions.

(66) Add chlorhydric acid, until the reaction becomes distinctly acid, and if a precipitate forms, wash it thoroughly with *cold* water upon a filter, until the filtrate is no longer acid to test paper. Observe whether sulphy- dric acid is given off.

If a precipitate is formed with chlorhydric acid, filter and test the filtrate according to page 101 (VI).

910:42

(67) *If no precipitate is formed pass to page 101 (VI).*

(a) If the precipitate is *white*, it may consist of:

PLUMBIC CHLORIDE. Boil a little of it with water, and test one portion of the solution for LEAD according to (63), and test another portion for CHLORINE by adding nitric acid and argentic nitrate;

PLUMBIC SULPHATE. Test according to (134);

ARGENTIC CHLORIDE. Test according to (135).

(b) If the precipitate is *colored* it may contain

The SULPHIDES of ARSENIC, ANTIMONY and TIN
Test, according to page 103, VIII.

SULPHUR may be precipitated, accompanied by a disengagement of sulphydric acid. The precipitated sulphur can be recognized by its appearance, and its insolubility in aqua regia.

V.—CHLORHYDRIC ACID TEST FOR ACIDS.

See also Silicic Acid (60).

If a solution is acid, many of the following tests, particularly (69), (70), and (71) can be applied, by simply heating the solution.

CARBONIC ACID is only present in alkaline solutions. It (68) is evolved with effervescence, when an acid is added, until the solution has an acid reaction. Hold a drop of lime-water on the end of a glass rod in the tube; if CARBONIC ACID, CO_2 , is present, a *white precipitate* forms. (See page 67, foot-note.) The same test can be applied to solid carbonates. See also the sulphuric acid test (40).

The following acids need only be looked for when an odor can be perceived after heating the solution, or after adding chlorhydric acid and heating.

CYANHYDRIC ACID, in its soluble combinations with (69) most metals, is set free by chlorhydric acid. It can be

recognized by its smell. See also the prussian blue test (127).

SULPHYDRIC ACID is evolved from alkaline solutions (70) (often with effervescence), on the addition of chlorhydric acid, when the solution is heated. It can be recognized by its smell, and by the lead paper test. (See the sulphuric acid test (43) and the argentic nitrate test (125).

SULPHUROUS ACID is evolved from alkaline, neutral or (71) slightly acid solutions, on the addition of chlorhydric acid. Mix a little potassic ferricyanide and ferric chloride, and hold a drop of the mixture on the end of a glass rod in the tube, after chlorhydric acid has been added, and the tube heated. If a *blue* color appears, **SULPHUROUS ACID**, SO_2 , is present. If sulphydric acid is present, add sufficient plumbic acetate to precipitate it, before performing the test. See the sulphuric acid test (44).

VI. SULPHYDRIC ACID TEST.

Metals in acid solutions.

- (72) Add sulphydric acid solution to a small quantity of the solution to be tested, and warm gently. In case metals of Group VI are to be tested for, it is better to pass sulphydric acid gas into the dilute solution, made acid with chlorhydric acid. The total precipitation of the metals of Group VI is frequently only effected after one or two days.

If no precipitate forms, no metals of Groups V and VI are present. Pass to page 109, X.

- (73) If a precipitate forms, observe the color. It may consist of the **SULPHIDES OF LEAD**, PbS ; **BISMUTH**, Bi_2S_3 ; **COPPER**, CuS ; **MERCURY**, HgS ; and **GOLD**, Au_2S_3 , when it is *black*; **ARSENIC**, As_2S_3 ; **TIN (BISULPHIDE)**, SnS_2 , *yellow*; **TIN (MONOSULPHIDE)**, SnS , *brown*; **ANTIMONY**, Sb_2S_3 or Sb_2S_5 , *orange*. The presence of a *black sulphide*

hides the color of the other sulphides, so that all may be present when the precipitate is black.

If only a light, fine, white precipitate, which is not destroyed by acids, is formed, it consists of sulphur, and is frequently due to the presence of a ferric salt in the solution. In case only sulphur is precipitated, pass to page 108, X.

- (74) If a precipitate forms in a small portion of the solution, a sufficient quantity for use in all the succeeding tests for metals must be treated with sulphydric acid until the metals of Groups V and VI are completely precipitated as sulphides; and the precipitate thus obtained must be washed on a filter quickly, with warm water containing sulphydric acid, until the addition of ammoniac hydrate to the filtrate ceases to produce a precipitate, and it must then be treated according to (75).

Test the filtrate for metals of Groups IV, III, II and I. (See page 108, X, &c.)

VII. SOLUBILITY OF THE SULPHYDRIC ACID PRECIPITATE IN AMMONIC SULPHIDE.

- (75) Add ammoniac sulphide to a small quantity of the precipitated sulphides (72), and warm gently.

If the precipitate dissolves entirely, it consists of the sulphides of metals of Group VI. Those of Group V are absent. Test the remainder of the precipitate according to (78) (a).

- (76) If a part of the precipitate does not dissolve, add four or five parts of water, and separate the solution by filtration from the undissolved precipitate.

The part of the precipitate which is insoluble in ammoniac sulphide, after being carefully washed, must be tested for sulphides of metals of Group V. (See page 105, IX.)

- (77) The ammoniac sulphide solution, obtained in (76), may contain metals of Group VI. Add to it gradually dilute

chlorhydric acid, until the solution becomes acid, and observe the color and general appearance of the precipitate which is produced. It is well to boil the liquid after the formation of a precipitate.

If only a fine white precipitate forms, which remains a long time in suspension in the liquid, even after boiling, it consists of sulphur, and metals of Group VI are absent, and the tests described in VIII can be omitted.

A flocculent precipitate, or one that becomes so on boiling, indicates the presence of metals of Group VI, and the color of the precipitate shows what metals predominate. Pass to the following tests :—

VIII. SEPARATION OF METALS OF GROUP VI.

- (78) If the test p. 102, VII, has shown the presence of metals of Group VI, and if the precipitate with sulphydric acid (73) was not entirely soluble in ammoniac sulphide, the whole of that precipitate must be treated two or three times with ammoniac sulphide, as directed (75), (76) and (77); and the sulphides of Group VI must be precipitated from the solution, and, after careful washing, must be dried at 100°.
- (a) Were the sulphides, precipitated by sulphydric acid, wholly soluble in ammoniac sulphide (see 75) it is sufficient to wash and dry the portion of the precipitate (73) which was not treated with ammoniac sulphide, and to use it for (79).*

* By heating the dry precipitate in a glass tube, or with less accuracy, by heating it before it is dry, on a bit of glass or porcelain, an approximate test may be made (see 13); and in case arsenic sulphide alone is indicated by the complete volatility of the precipitate, this test is conclusive, and the remaining tests in the separation of metals of Group V may be omitted. The test has little value, except when the pure yellow color of the precipitate gives rise to the suspicion that only arsenic sulphide is present.

(79) Detach the precipitate (78) from the filter (or, if this is impossible, cut the portion of the filter, to which it adheres, in small pieces), and mix it by rubbing in a mortar with one part of sodic carbonate and one part of sodic nitrate. Fuse two parts of sodic nitrate in a small porcelain crucible, but do not heat to the point at which decomposition with evolution of gas ensues. Transfer the mixture of the precipitated sulphides and the sodic carbonate and nitrate in *small portions* to the fused nitrate in the crucible; immediate oxidation takes place, and as soon as the sulphides are burnt white, pour the fused mass out on a clean bit of porcelain, place it in a test tube, and digest it with a small quantity of cold water until it entirely dissolves, or until only a flocculent precipitate remains insoluble.

(80) If a portion remains insoluble, it may contain TIN, as STANNIC ACID, SnO_2 ; and ANTIMONY, as ANTIMONIATE OF SODIUM, Na_3SbO_4 . It should be filtered, and the filtrate preserved. The soluble portion may contain ARSENIC, as ARSENIATE OF SODIUM, Na_3AsO_4 . The precipitate on the filter should be washed with a mixture of equal parts of water and alcohol, and the washings should be thrown away.

ARSENIC. To test the first filtrate (80), (or the solution, if (81) the fused mass was entirely soluble,) for arsenic, render a portion of it slightly acid with acetic acid, and add nitrate of silver. If a *red* precipitate forms, it consists of ARSENIATE OF SILVER, Ag_3AsO_4 . Additional test: Another portion of the above mentioned solution may be used for *Marsh's* test, with perfectly pure zinc and sulphuric acid (see page 56), and if metallic flakes are found in the apparatus some time after the solution has been poured into it, they should be tested for TIN by dissolving them in chlorhydric acid and adding mercuric chloride (see

Separation and detection of
Arsenic, Tin and Antimony.

Collect the pptd. sulphides obtained in 78, page 103, Qual. Anal. on a filter wash with hot water, free the ppt. as completely as possible from adhering water by passing the filter and its contents between several thicknesses of filter paper, remove the ppt from the filter and heat it with concentrated HCl .

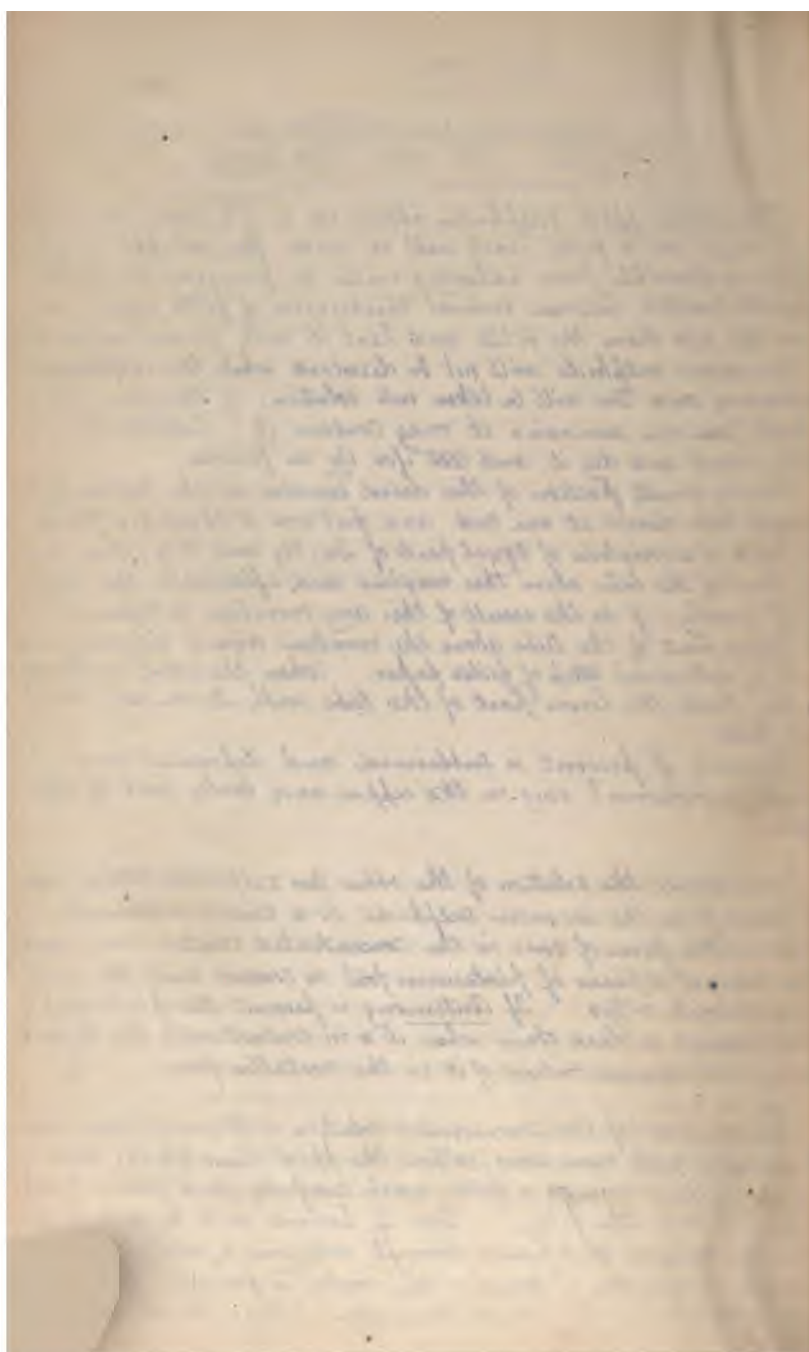
Arsenous sulphide will not be dissolved, while the sulphides antimony and tin will be taken into solution; if therefore an soluble residue remains it may contain As . Collect it on filter, wash and dry it, and test for As as follows.

Put a small portion of this dried residue at the bottom of a small tube closed at one end, and put over it about six times its bulk of a mixture of equal parts of H_2CO_3 and HCl ; then heat the part of the tube above this mixture and afterwards the mixture itself, gently; if as the result of this any moisture is deposited the upper part of the tube above the mixture wipe it out carefully with a rolled-up strip of filter paper. When the whole is thoroughly dry heat the lower part of the tube with its contents to a red heat.

Arsenic if present is sublimed and deposited as a black or brownish ring in the upper and cooler part of the tube.

Concentrate the solution of the other two sulphides, that is filtered from the Arsenous sulphide, in a small evaporator, put a little piece of zinc in the concentrated solution, and bend the edge of a piece of platinum foil in contact with the zinc for a minute or two. If Antimony is present the platinum will receive a black stain where it is in contact with the liquid owing to the precipitation of it in the metallic form.

Put the rest of the concentrated solution of the sulphides in an evaporator with more zinc, collect the pptd black flakes that appear after a time on a filter, wash carefully and pour a little conc. HCl over the filter. Tin if present will be dissolved and the solution that passes through will give a white ppt. or perhaps a gray one if much of the metal is present, with arsenic chloride. (See Crafts Qual. Anal. page 52 at bottom)



83). (This precaution is necessary, because, if nitrate of soda is heated too strongly, caustic soda is formed, which dissolves STANNIC ACID, SnO_2 ; hence the tin may be found in the solution, and not in the insoluble part of the fused mass (**79**)).

ANTIMONY. Dissolve the portion of the fused mass (**79**) (**82**) which was insoluble in cold water, by pouring hot chlorhydric acid on the filter; dilute the solution with water, and place in it, in contact with each other, a piece of pure zinc and a slip of platinum foil.

If ANTIMONY is present, a *black stain* of the reduced metal quickly appears on the platinum.

TIN. After allowing the action of the zinc, described in the (**83**) last paragraph, to continue for a few minutes, if metallic flakes are precipitated, wash them with water by decantation, and dissolve them in a very small quantity of hot concentrated chlorhydric acid. Dilute the solution, and add to it a solution of mercuric chloride. If a white precipitate of MERCUROUS CHLORIDE, Hg_2Cl_2 , forms (frequently not till after the lapse of a few minutes), TIN is present.

GOLD. (The analytical chemist usually knows whether it is (**84**) necessary to test for gold or not. In ordinary analyses its presence would be improbable.) When gold is present, it remains on the filter after the treatment with hot chlorhydric acid (described in **82**). Pour upon the filter a mixture of nitric and chlorhydric acids. Evaporate the solution nearly to dryness, dilute with water, and add ferrous sulphate solution. The formation of a *brown* or *purple* precipitate of METALLIC GOLD, either immediately or after heating, indicates the presence of the metal.

IX. SEPARATION OF METALS OF GROUP V. (SECTION II.)

(**85**) If a portion or the whole of the precipitate obtained

with sulphydric acid is insoluble in ammoniac sulphide (see 76), free it by careful washing from the liquid in which it was formed, or from the ammoniac sulphide which was used to dissolve the soluble portion, place it in a porcelain dish, pour upon it pure concentrated nitric acid, and heat it gently, if red fumes are given off, until they cease. In any case, complete the operation by adding a little water, and boiling the contents of the dish for a few minutes.

If no part of the precipitate, or if only yellow particles of sulphur remain insoluble, mercury is absent. Pass to (87).

In this test, when the liquid holding sulphur in suspension is boiled, the sulphur melts, and may enclose particles of black sulphides, which then become very difficult to dissolve, and the appearance of the sulphur may, in such a case, lead to an erroneous conclusion that mercuric sulphide is present. It is for this reason that the precipitate is oxydized with strong nitric acid, as far as possible at a temperature below its boiling point, before the sulphides are finally boiled with a somewhat weaker acid. The same cause makes the confirmatory test for mercury with stannous chloride (86) necessary, when there appears to be an insoluble black sulphide.

MERCURY. If a black sulphide, HgS , remains insoluble (86) after the above treatment (85), a MERCURIC SALT is probably present. Confirmatory test: Boil the black insoluble sulphide with chlorhydric acid and a little potassic chlorate in a porcelain dish, and evaporate until the greater part of the acid is volatilized; dilute with water (it is not necessary to filter), and add stannous chloride. A white precipitate of MERCUROUS CHLORIDE, Hg_2Cl_2 , is formed, if mercury is present.

If mercury is present, dilute a few drops of the nitric

acid solution of the sulphides (85) with water, and add sulphydric acid. If a *black* precipitate is formed, other metals of Group V are present; pass to (87); *otherwise they are absent*; pass to page 108, X.

LEAD. Add a few drops of strong sulphuric acid to a small (87) portion of the nitric acid solution of the sulphides (85), and evaporate until *dense, white* fumes of sulphuric acid appear, and dilute with a considerable quantity of water. If a *white* precipitate forms, it consists of SULPHATE OF LEAD, PbSO_4 .

The test can be made more delicate by adding an equal bulk of alcohol to the solution after it has been diluted with water. If lead is discovered, treat the whole of the nitric acid solution of the sulphides in the same way; filter, and use the filtrate for (88).

BISMUTH. Add ammoniac hydrate to alkaline reaction, to (88) the nitric acid solution of the sulphides, or to the filtrate from the lead precipitate, if lead was present. If bismuth is present, it is precipitated as the HYDRATE OF BISMUTH, $\text{Bi}(\text{HO})_3$, *white*. If bismuth is present filter, and use the filtrate for (89).

COPPER. If copper is present, it is dissolved by the ammoniac (89) hydrate, and imparts a *blue* color to the solution.



METALS OF GROUPS IV & III.

X.—AMMONIC SULPHIDE TEST.

- (90) To a small portion of the filtrate from the precipitate produced by sulphydric acid, or to a portion of the original solution, if no precipitate is produced in it by sulphydric acid, add sufficient ammoniac hydrate (free from carbonate) to make the reaction alkaline, and then, whether a precipitate is formed or not, add ammoniac sulphide. If the solution contains no chlorhydric acid, it is necessary to add a small quantity before neutralizing with ammoniac hydrate.

If a *black* precipitate is formed, it may contain the SULPHIDES OF NICKEL, NiS ; COBALT, CoS ; IRON, FeS ; MANGANESE, MnS ; and ZINC, ZnS ; and the HYDRATES OF ALUMINIUM, $\text{Al}_2(\text{HO})_6$, and CHROMIUM, $\text{Cr}_2(\text{HO})_6$. If the precipitate is *white* or *flesh-colored*, it can only consist of the SULPHIDES OF MANGANESE AND ZINC, and the HYDRATES OF ALUMINIUM AND CHROMIUM. *In this latter*

case omit (92), (93), (94), and ~~95~~ 97

If no precipitate is formed no members of Groups III and IV are present. Pass to page 115, XII.

- (91) If a precipitate was formed in the above test, (90), treat a considerable quantity of the solution in the same way, heat the liquid, filter as quickly as possible, and wash immediately with boiling water, until the filtrate has no longer an alkaline reaction.

The filtrate must be tested according to page 115, XII.

The detection of Cobalt and Nickel.

If the filtrate from the ppt. by (NH₄)₂S obtained in 91 has a deep brown color, nickel may safely be put down as present in the substance analyzed; of course in that case an insoluble black residue will be left on treatment of this ppt. with dilute HCl. a blue bead can be obtained with this residue, as described in 98 (Casts). Cobalt is present; if a bead is not obtained, but only a light-brown or nearly colorless one, cobalt is absent.

If all these reactions are obtained in an unmistakable manner there will be no necessity of making further tests for these two elements.

If however a black insoluble residue is obtained in 92, and there was no brown filtrate in 91, or if using for Cobalt with the bead, the color was so dark brown, owing to the presence of a large proportion of Nickel, that the blue Cobalt color might have been masked, the following method must be followed for separating the two substances.

Transfer the insoluble sulphides obtained in 92 to an evaporator, add concentrated HCl and a drop of conc. H₂O₂, boil (in the hood) till the liquid is nearly all evaporated away, add NH_4OH to the residue till it is strongly alkaline, and then add a quantity of sodium hypochlorite nearly equal to the bulk of the liquid in the dish; boil the mixture a few minutes, filter, wash the ppt on the filter, return it to the evaporator, add ammonia, boil again, filter if a residue remains, and, finally, pass through the filtrate, a black ppt or a deep brown coloration of the liquid indicates Nickel.

Test the residue from which this liquid was filtered for Cobalt with the brown bead as directed in 98 page 100.



SULPHIDES INSOLUBLE IN DILUTE CHLOR-
HYDRIC ACID.

- 92)** Add to the precipitate (**91**) cold, dilute chlorhydric acid; if a *black* residue is insoluble, it consists of the SULPHIDE OF NICKEL OR COBALT. Filter and examine the residue on the filter according to (**93**) and (**94**).

*The filtrate must be tested according to (**95**).*

*If the precipitate dissolves entirely, or if only a white residue is left, no NICKEL or COBALT are present. Pass to (**95**).*

OBALT. Dissolve a portion of the precipitate, which **93)** proved to be insoluble in cold, dilute chlorhydric acid, in the borax bead. Examine the color of the bead, after flattening it while hot between the stopper of a bottle and the bead of the blowpipe. If head the is *blue*, COBALT is present. If the bead is *brown*, NICKEL is present in large quantity. The bead is *blue* even when more nickel than cobalt is present. To test for traces of cobalt in a nickel-bead, shake the hot bead from the wire, heat it two or three minutes on charcoal in a good reducing flame, remove it from the charcoal, melt it on the platinum wire in the reducing flame, and flatten it as before. Even if only *traces* of cobalt are present, the bead will be colored blue.

NICKEL. When nickel is present and when it is nearly free **94)** from cobalt, it can be discovered by the *brown* color, which it imparts to the borax-bead, and the test is conclusive.

To discover small quantities of nickel in the presence of considerable quantities of cobalt, dissolve a large portion of the precipitate, insoluble in cold, dilute chlorhydric acid (see **92**), by heating a small quantity of dilute sul-

phuric acid,* adding to it a few drops of strong nitric acid, and pouring the hot mixture on the precipitate; dilute with a little water, add a little potassic ferricyanide, and then add to the liquid an equal bulk of ammoniacal hydrate, and finally add baric chloride,* until the color of the precipitate has become very much lighter. Allow the precipitate to settle for a few minutes, and filter. Add to the filtrate ammoniacal sulphide. If NICKEL is present, a black color or a black precipitate is formed.

SULPHIDES, SOLUBLE IN DILUTE. CHLORHYDRIC ACID.

- (95) Boil in an evaporating dish the solution, which was obtained by the treatment of the ammoniacal sulphide precipitate with cold, dilute chlorhydric acid (see 92), until the smell of sulphydric acid has entirely disappeared, add a few drops of strong nitric acid, and boil a minute longer, and filter if a precipitate of sulphur has formed.

If oxalic, phosphoric, and boracic acids are not known to be absent pass to page 112, XI.

- (96) Add sodic hydrate to the solution obtained in (95) until the reaction becomes very strongly alkaline, dilute with water and boil for a few minutes.

If a precipitate forms immediately or after boiling, test it according to (97), &c.; it may contain the HYDRATES of IRON, $\text{Fe}_2(\text{HO})_6$, MANGANESE, $\text{Mn}(\text{HO})_2$, and CHROMIUM, $\text{Cr}_2(\text{HO})_6$. In this case filter and test the filtrate, which may contain ZINC and ALUMINIUM, according to (100), (101.)

* The only object of adding sulphuric acid and baric chloride is to form a precipitate of baric chloride in the solution, which carries down with it the very light precipitate of cobaltic ferricyanide and renders the filtration easier. The addition of these may be omitted, and it is then usually necessary to reject the first portion of the filtrate and only to use that which comes through clear.

If no precipitate forms, pass to (100). The solution may contain ZINC and ALUMINIUM.

IRON. Put a small quantity of the precipitate, (96) on a (97) watch-glass, dissolve it in a single drop of dilute chlorhydric acid, dilute with water, and add potassic sulphocyanate. A *red* color indicates the presence of IRON.

MANGANESE. Fuse a little carbonate of soda in the loop of (98) a platinum wire, take up some of the precipitate, obtained in (96), on the bead thus formed, and heat it in the oxydizing flame of the blowpipe or Bunsen's lamp. The appearance of a *green* color in the bead is proof of the presence of MANGANESE.

CHROMIUM. Fuse another portion of the precipitate, obtained in (96), on platinum foil with sodic carbonate and sodic nitrate, boil the fused mass with a little water, either on the foil or in a small test-tube; if a precipitate is formed, filter through a very small filter, and acidify with acetic acid. The appearance of a *yellow* color in the solution is proof of the presence of CHROMIUM.

Acetate of lead may also be added to the solution, acidified with acetic acid, and it then takes a deeper *yellow* color and becomes turbid, or a *yellow* precipitate is formed if chromium is present. The formation of a white precipitate under these circumstances only indicates that the carbonate of sodium employed contained sulphate of sodium as an impurity.

ZINC. Add sulphydric acid to a portion of the solution in solution (100) dic hydrate, obtained in (96), after it has been filtered, if a precipitate was formed. If a *white flocculent* precipitate forms, it consists of SULPHIDE OF ZINC, ZnS .

If chromium was discovered in (99), zinc may also be present in the precipitate, obtained in (96). Therefore, in that case, dissolve a portion of the precipitate, by boiling with a very little dilute chlorhydric acid, add so-

dic hydrate until the reaction is alkaline, acidify with acetic acid and add sulphydric acid. A *white flocculent* precipitate consists of SULPHIDE OF ZINC, ZnS .

ALUMINIUM. Add to another portion of the sodic hydrate (101) solution, obtained in (96), chlorhydric acid, until the reaction becomes acid, and then ammoniac hydrate, until it becomes alkaline, and boil. If a *white flocculent* precipitate forms, it consists of ALUMINIC HYDRATE, $\text{Al}_2(\text{HO})_6$. This precipitate is at first gelatinous, and it may easily escape notice; it is therefore best to set the test-tube aside, and to wait a quarter of an hour for the precipitate to settle.

XI.—AMMONIAC SULPHIDE TEST IN CASE PHOSPHORIC, OXALIC AND BORACIC ACIDS ARE PRESENT.

If PHOSPHORIC, OXALIC OR BORACIC ACID was present in the original solution, these acids together with the metals of Group II may be contained, wholly or in part, in the solution obtained in (95); for the acids would be precipitated with any of the metals of Groups II, III and IV, during the treatment with ammoniac hydrate and ammoniac sulphide in (90), and the precipitates would be dissolved during the treatment with dilute chlorhydric acid in (92), and consequently the metals and acids might be contained in the solution (95). The solution (95) must be tested for phosphoric, oxalic and boracic acids and freed from them before the ordinary course of analysis can be proceeded with.

PHOSPHORIC ACID. Use the test (123) with a small portion of the solution obtained in (95).

If phosphoric acid is present, use (104) and the succeeding tests.

If phosphoric acid is absent use (96) and the succeeding tests.

In either case first perform the operation described in the next paragraph.

(102) Before testing for OXALIC AND BORACIC ACIDS it is necessary to set them free from their combinations with the metals of Groups II, III, and IV. To effect this end add sodic carbonate to a small quantity of the solution obtained in (95) until the reaction becomes strongly alkaline and boil for a few minutes and filter. The metals are precipitated with the exception of a portion of the aluminium and the acids remain in the solution.

[If no precipitate is formed with sodic carbonate it is unnecessary to test further for these metals or acids, as in that case they cannot be present in the solution (95)].

OXALIC ACID. Test a portion of the filtrate obtained after boiling with sodic carbonate for oxalic acid according to (121).

BORACIC ACID. Test another portion of the same filtrate for boracic acid according to (124).

If one or both these acids are found, the whole of the remainder of the solution obtained in (95) must be treated with sodic carbonate as in (102) and the precipitate thus obtained must be dissolved in dilute chlorhydric acid. The solution must be used for (105) and the succeeding tests, if phosphoric acid was discovered.

If phosphoric acid is absent, the solution must be used for (96) and the succeeding tests.

ALUMINIUM. If it was found necessary to treat the solution (103) obtained in (95) with sodic carbonate according to (102) the filtrate from the precipitate produced by sodic carbonate may contain a portion of the aluminium, add to the filtrate in (102) chlorhydric acid until the reaction becomes acid, and then ammoniac hydrate until the reaction becomes alkaline. If a *white, flocculent* precipitate forms, immediately, or after long standing, it contains aluminium.

METALS OF GROUP II.

XII. DETECTION OF BARIUM, CALCIUM AND MAGNESIUM.

(108) To a small portion of the solution, to which the previous tests have been applied, or to a solution which has been found to contain no metals of the higher groups, add ammoniac hydrate until the reaction becomes alkaline, and then ammoniac carbonate, and heat gently. (If the solution does not already contain ammoniac chloride, this also must be added, to prevent the precipitation of magnesium.)

If a white precipitate forms, it can only consist of BARIUM CARBONATE, BaCO_3 , and CALCIC CARBONATE, CaCO_3 .

If no precipitate forms, BARIUM and CALCIUM are absent; pass to (112).

If a precipitate was formed with ammoniac carbonate, the whole of the solution must be treated as described above. Filter, wash the precipitate, and test the filtrate according to (112). Dissolve the precipitate by pouring a very little dilute chlorhydric acid on the filter, and use the solution thus obtained for (109)—(111).

BARIUM. To a small portion of the solution in chlorhydric (109) acid add a considerable quantity of calcic sulphate. If a precipitate forms, it consists of BARIUM SULPHATE, BaSO_4 .

(110) If BARIUM is discovered by means of calcic sulphate, it must be completely precipitated by dilute sulphuric acid in the remainder of the solution, and the filtrate so obtained tested for calcium. Therefore, in case BARIUM is present, add dilute sulphuric acid to the remainder of the solution, boil, filter, and use the filtrate for (111).

IRON. Add a few drops of potassic sulphocyanate to the solution obtained in (95); a *red* color indicates the presence of IRON.

(105) . Add to the remainder of the solution obtained in (95), if phosphoric acid alone is present, or to the solution obtained in (102), if oxalic or boracic acid is likewise present, ferric chloride, until a few drops, treated with ammoniac hydrate on a watch-glass, give a *yellow* and not a *white* precipitate; dilute largely with water, render the solution alkaline with ammoniac hydrate, and add a considerable excess; then add acetic acid, until the solution has a slight acid reaction, and boil for a few minutes in a flask.

The precipitate contains all the IRON, ALUMINIUM, CHROMIUM and PHOSPHORIC ACID present in the solution so treated. Test according to (106) and (107).

The filtrate contains the MANGANESE, ZINC, and probably part of the BARIUM, CALCIUM and MAGNESIUM, which were present in the original solution.

The operations described, page 108, X, and page 115, XII, must be repeated with this filtrate, omitting those which relate to the separation of NICKEL and COBALT, and the detection of IRON, CHROMIUM and ALUMINIUM.

CHROMIUM. Test a portion of the precipitate obtained (106) in (105), for chromium according to (99).

ALUMINIUM. Boil the remainder of the precipitate obtained (107) in (105), with sodic hydrate, and test the solution for aluminium according to (101).

(The method of precipitation by boiling the acetic acid solution used in (105), can be used in all cases for the separation of aluminium, chromium and iron (ferric salts) from the metals of all other groups, except the Groups V and VI, and it is preferable to any other, but it demands more skill in manipulation.)

METALS OF GROUP II.

XII. DETECTION OF BARIUM, CALCIUM AND MAGNESIUM.

- (108) To a small portion of the solution, to which the previous tests have been applied, or to a solution which has been found to contain no metals of the higher groups, add ammoniac hydrate until the reaction becomes alkaline, and then ammoniac carbonate, and heat gently. (If the solution does not already contain ammoniac chloride, this also must be added, to prevent the precipitation of magnesium.)

If a white precipitate forms, it can only consist of BARIUM CARBONATE, BaCO_3 , and CALCIC CARBONATE, CaCO_3 .

If no precipitate forms, BARIUM and CALCIUM are absent; pass to (112).

If a precipitate was formed with ammoniac carbonate, the whole of the solution must be treated as described above. Filter, wash the precipitate, and test the filtrate according to (112). Dissolve the precipitate by pouring a very little dilute chlorhydric acid on the filter, and use the solution thus obtained for (109)—(111).

BARIUM. To a small portion of the solution in chlorhydric (109) acid add a considerable quantity of calcic sulphate. If a precipitate forms, it consists of BARIUM SULPHATE, BaSO_4 .

- (110) If BARIUM is discovered by means of calcic sulphate, it must be completely precipitated by dilute sulphuric acid in the remainder of the solution, and the filtrate so obtained tested for calcium. Therefore, in case BARIUM is present, add dilute sulphuric acid to the remainder of the solution, boil, filter, and use the filtrate for (111).

If no barium is present, use the remainder of the solution obtained in (108) for (111).

CALCIUM. Render the solution alkaline by the addition of (111) ammoniac hydrate, and add ammoniac oxalate. If a *white* precipitate forms, it consists of CALCIC OXALATE, CaC_2O_4 .

If calcium is present, add ammoniac oxalate until a few drops of the liquid passed through a filter give no further precipitation with that reagent, filter, and use the filtrate for the magnesium test.

If no calcium is present, use the solution to which ammoniac oxalate was added, without filtering it, for the magnesium test.

MAGNESIUM. Add to the solution sodic phosphate. If a (112) *white* precipitate forms (frequently only after the lapse of some minutes), it consists of the PHOSPHATE OF MAGNESIA AND AMMONIA, MgNH_4PO_4 .

METALS OF GROUP I.

XIII. DETECTION OF SODIUM, POTASSIUM AND AMMONIUM.

Before testing for sodium and potassium, precipitate the metals of Groups V and VI which are present, with sulphydric acid, and precipitate those of Groups II, III and IV with a mixture of ammonio carbonate and sulphide, and use the solution, freed from those metals, for the following tests (113) and (114).

SODIUM. Evaporate the solution to dryness, and drive off (113) ammonia salts, if any are present, by heat. Sodium, if present, can be distinguished by the *yellow* color, which a small quantity of the solid residue held in the flame of a gas or alcohol lamp imparts to it. See page 25 (36).

Frequently sodium can be detected in a solution without evaporation, by dipping a platinum wire in the solution, and then holding it in the flame.

POTASSIUM can be recognized by the violet color which it (114) imparts to the flame. The solid residue obtained in (113) can be tested for potassium in the same way that it is tested for sodium.

If sodium is also present, its greater coloring power will obscure the potassium flame, but by looking through a piece of blue glass at the flame, the violet color can be distinguished even when sodium is present. (The color of potassium flame is almost the same as that of the platinum wire, while the sodium flame is much more blue, if it is not excluded entirely by the glass.)

AMMONIUM. Add to a portion of the original solution a few (115) drops of sodic hydrate, and heat.

AMMONIA can be recognized by its smell, or by holding a piece of moistened turmeric paper or red litmus paper at the mouth of the test tube, taking care not to let it touch the sides, which may be moistened with sodic hydrate. The turmeric paper will be turned brown, and the litmus paper blue, if ammonia is present.

TESTS FOR ACIDS.

It is usual to take a fresh quantity of the solution to test for acids, and generally the tests for metals precede those for acids, in order that information gained by the first series of experiments may point out the most convenient way of detecting the acids. Silicic acid, however, is always first precipitated from a solution as directed, page 95 (60), and the acids noticed under the chlorhydric acid test (page 100, V) are to be looked for while that test is applied to the detection of the metals. Phosphoric, oxalic and boracic acids interfere with the ordinary methods of testing for the metals of Groups II, III, and IV, and consequently these acids must be tested for as directed, page 112, XI, during the application of the tests for the metals of those groups.

In all other cases, when the presence of any metals or acids in the solution interferes with the performance of the test for an acid, directions are given under the head of each acid for removing them.

It is obvious that metals and acids, which precipitate each other, can not be present together in a solution, and that when certain metals have been found, the number of acids which it becomes necessary to look for is restricted within limits determined by this consideration. Therefore the knowledge already acquired of the composition of a solution must be brought to bear upon the problem of testing for acids.

First, the reaction, which the solution gives with test-paper, must be considered, and then the tables IV and V must be consulted to ascertain what acids can exist in a solution, possessing

the observed reaction, together with the metals which have been discovered.

For instance, if a solution contains *lead*, and is neutral or nearly neutral, the only acids which can be present in it are acetic, chlorhydric, chloric and nitric. If the reaction is strongly acid the solution may contain all the acids except sulphuric, sulphydric, ferro- and ferricyanhydric. Moreover, the solution can not contain a large quantity of lead and chlorhydric acid at the same time, because the chloride of lead is only soluble in 135 parts of water.

ACIDS OF GROUP I.

ARSENIOUS AND ARSENIC ACIDS are always detected (116) by the sulphydric acid test, in searching for the metals.

When these acids are discovered, they must always be precipitated by sulphydric acid, before testing further.

CHROMIC ACID cannot be present in a solution to which (117) SULPHYDRIC ACID OR AMMONIC SULPHIDE has been added (see page 60). If chromic acid is present in a solution, it must be contained in the precipitate obtained with baric chloride (118), and can be detected by heating a small portion of the precipitate in the borax bead. If **CHROMIC ACID**, H_2CrO_4 , is present, the bead will be colored *green*. Chromic acid can often be recognized by the *yellow* color which it imparts to solutions which contain it, and by the yellow precipitate, $PbCrO_4$, which is obtained by adding plumbic acetate to the neutral or slightly acid solution.

XIV.—BARIC CHLORIDE TEST.

Baric nitrate and nitric acid should be used instead of baric chloride and chlorhydric acid, when lead, silver or mercurous salts have been discovered in the solution.

ACIDS OF GROUP I. Put a piece of litmus paper in the (118) solution, and if the reaction is acid, add ammoniac hydrate drop by drop, until it becomes slightly alkaline. If a precipitate is formed in consequence, add dilute chlorhydric acid only in sufficient quantity to dissolve it. Add baric chloride, and if a precipitate forms, it indicates the presence of ARSENIOUS, ARSENIC, CHROMIC, SULPHURIC, SULPHUROUS, OXALIC, FLUORHYDRIC, PHOSPHORIC, BORACIC, CARBONIC, AND SILICIC ACIDS. (Silicic acid cannot be present after the operation, (60), has been performed).

If these acids are absent, pass to the argentic nitrate test (page 123, XVIII).

(The baric chloride test is of little value except when the substance is soluble in water with a neutral or slightly alkaline reaction, or in case the reaction is acid, when the metals of Groups II, III, IV, and V are absent. The following special tests are more accurate.)

SULPHURIC ACID. Acidify (if the solution is not already (119) acid) with dilute chlorhydric acid, in considerable excess, and add baric chloride as long as a precipitate continues to form. (Use nitric acid and baric nitrate, if chlorhydric acid produces a precipitate.) If sulphuric acid is present, it is precipitated as BARIC SULPHATE, BaSO_4 , *fine white powder*. The solution must not be heated, when it is intended to use the filtrate for the next test.

SULPHUROUS ACID. To the filtrate from the precipitate (120) produced by baric chloride, or to the acid solution to which baric chloride has been added without producing a precipitate, add potassic bichromate, and boil. If a precipitate forms, it consists of BARIC SULPHATE, BaSO_4 , produced by the oxydation of SULPHUROUS ACID, H_2SO_3 , contained in the solution. Usually the CHLORHYDRIC ACID TEST (71) is more convenient and sufficiently accurate.

XV.—CALCIC SULPHATE TEST.

If the solution contains metals, which are precipitated by sulphydric or sulphuric acids, they must first be removed by adding a slight excess of those precipitants and filtering.

OXALIC ACID, if the solution is alkaline, add acetic acid (121) until the reaction becomes acid; if it is acid, add sodic hydrate until the reaction becomes alkaline, and then add acetic acid until it becomes acid, and test as below for oxalic acid. If a precipitate forms and does not dissolve in the acetic acid, add to the original solution a considerable excess of sodic carbonate, boil, filter, add to the filtrate acetic acid, until its reaction becomes acid, and test as follows for oxalic acid: Add calcic sulphate in considerable quantity. If a precipitate forms, it consists of CALCIC OXALATE, CaC_2O_4 , *white powder*. Fluorhydric acid is the only other acid which precipitates calcic sulphate under these circumstances, and as the precipitate is *almost transparent* and *gelatinous*, it cannot easily be mistaken for that produced by oxalic acid.

FLUORHYDRIC ACID can only be present in alkaline (122) solutions in glass vessels. If there is reason to suspect the presence of this acid, add calcic chloride and ammoniac hydrate to the solution, and if a precipitate forms, collect it on a filter, and examine it for fluorine according to (42).

XVI.—AMMONIC MOLYBDATE TEST.

PHOSPHORIC ACID. Make the solution strongly acid (if (123) it is not so already) with nitric acid, and add a small portion of it to a considerable quantity of ammoniac molybdate solution.* If phosphoric acid is present, PHOSPHO-MOLYBDATE OF AMMONIUM, *yellow, crystalline*

* See foot-note, page 64.

powder, is precipitated. If the quantity of phosphoric acid in the solution is very small, the precipitate does not form until after several hours.

If sulphydric acid is present, it is necessary to heat the acid solution until it is expelled, before performing the test. See (130) a) for this test in the presence of ferro- or ferricyanhydric acid.

XVII.—TURMERIC-PAPER TEST.

BORACIC ACID.* Strongly acidify the solution (if it is not (124) already acid) with dilute chlorhydric acid, dip a piece of turmeric paper in it, and dry the paper, by holding it over the lamp flame, without charring it. If a *red* or *brownish-red* stain appears upon the paper, when it is dry, it is due to the presence of BORACIC ACID, H_3BO_3 .

ACIDS OF GROUP II.

XVIII.—ARGENTIC NITRATE TEST.

Acidify with nitric acid, if the solution is not already acid, and add argentic nitrate.

SULPHYDRIC ACID. If a *black* precipitate is formed, (125) it must contain ARGENTIC SULPHIDE, Ag_2S , showing that sulphydric acid was present in the solution. The precipitate may also contain ARGENTIC CHLORIDE, CYANIDE, FERRO- and FERRICYANIDE.

If a precipitate is formed which is not *black*, it can only be due to presence of CHLORHYDRIC and CYANHYDRIC, FERROCYNHYDRIC, and FERRICYANHYDRIC ACIDS.

If no precipitate is formed, none of the above acids are present. Pass to (131).

* See foot-note, page 65.

CHLORHYDRIC ACID (see also (46)). Acidify strongly (126) with concentrated nitric acid, add argentic nitrate in excess, shake thoroughly if a precipitate forms, and allow it to settle, decant the liquid, and pour on the precipitate strong nitric acid, and boil for five minutes; if a precipitate remains undissolved, it consists of ARGENTIC CHLORIDE, AgCl .

See (130) (a) for this test in the presence of ferro- or ferricyanhydric acid.

XIX.—PRUSSIAN BLUE TEST FOR CYANHYDRIC ACID.

CYANHYDRIC ACID. Add to the solution ferrous sulphate (127) and a few drops of ferric chloride, add sodic hydrate until a precipitate forms (unless the solution is alkaline and a precipitate forms without the addition of sodic hydrate), warm for a minute and add dilute chlorhydric acid, until the reaction becomes acid. The appearance of a *blue* precipitate or a *blue* color in the solution is evidence of the presence of cyanhydric acid.

See (130) (b) for this test in the presence of ferro- or ferricyanhydric acid.

XX.—FERRIC CHLORIDE TEST.

FERROCYANHYDRIC ACID. Add a little ferric chloride (128) to the acid solution. If ferrocyanhydric acid is present, a precipitate of PRUSSIAN BLUE, $\text{Fe}_4(\text{FeCy}_6)_3$, *deep blue*, is formed.

XXI.—FERROUS SULPHATE TEST.

FERRICYANHYDRIC ACID. Add a little ferrous sulphate (129) to the acid solution. If ferricyanhydric acid is present, a precipitate of TURNBULL'S BLUE, $\text{Fe}_3(\text{Fe}_2\text{Cy}_{12})$, *deep blue*, is formed.

(130) If ferro- or ferricyanhydric acid is present, before per

- (a) forming the tests for PHOSPHORIC ACID (123), and CHLORHYDRIC ACID (126), the following steps must be taken: Add dilute sulphuric acid, dilute with water if the solution is not dilute, add cupric sulphate, and finally add enough baric nitrate* to render the precipitate of a decidedly lighter color; heat almost to boiling, allow the precipitate to settle for a few minutes, filter, and use the filtrate for the tests (123) and (126). In case the test (47) for nitric acid is to be used take the same preliminary steps but using baric chloride in place of baric nitrate.
- (130) If ferro- or ferricyanhydric acid is present, the test for
- (b) CYANHYDRIC ACID (127) is to be modified in the following manner: Dilute with water if the solution is not dilute, add dilute sulphuric acid, then add, according as ferro- or ferricyanhydric acid is present, ferric chloride or ferrous sulphate, or both, in sufficient quantity to precipitate the ferro- or ferricyanhydric acid or both acids; finally, add baric chloride* until the color of the precipitate has become decidedly lighter; shake thoroughly, allow the precipitate to settle for a few minutes, and filter. Add to a portion of the filtrate sodic hydrate until a precipitate forms, warm gently, and add dilute chlorhydric acid until the solution becomes acid. The appearance of a blue precipitate, or of a *blue color*, is evidence of the presence of CYANHYDRIC ACID.

* Sulphuric acid and baric nitrate or chloride are only added in order to produce a heavy precipitate of baric sulphate, which carries down with it the lighter particles of the other precipitates, and renders the filtration easier.

ACIDS OF GROUP III.

ACIDS WHICH ARE PRECIPITATED BY NO
METALS.

CHLORIC ACID. See sulphuric acid test (45).
(131)

NITRIC ACID. See sulphuric acid test (47).
(132)

ACETIC ACID. See sulphuric acid test (48).
(133)

CLASS III.

SUBSTANCES WHICH ARE INSOLUBLE IN WATER AND IN ACIDS.

(See page 91.)

XXII.—The only substances which are insoluble after the treatment described, on page 91, are the following :

PLUMBIC SULPHATE (not absolutely insoluble in acids).

ARGENTIC CHLORIDE (slightly soluble in chlorhydric acid).

SULPHUR.

CARBON.

BARIC SULPHATE, SILICIC ACID and many SILICATES, and some OXIDES.

XXIII.—SOLUTION IN AMMONIC ACETATE AND POTASSIC CYANIDE.

PLUMBIC SULPHATE. Boil a portion of the substance (134) with ammonic acetate, and test the solution (after filtration, if necessary) with ammonic sulphide. If LEAD is present, a *black* color or a *black* precipitate is formed.

Test the solution also for SULPHURIC ACID, according to (119).

If lead is discovered, repeat the treatment with ammonic acetate, until no more lead is dissolved.

ARGENTIC CHLORIDE. Digest a portion of the substance (135) free from plumbic sulphate, with potassic cyanide, warm, (unless it *blackens* by warming), and test the solution (after filtration, if necessary) with ammonic sulphide. A *black* precipitate indicates the presence of SILVER. If a *black* precipitate forms, wash it, dissolve it

in strong nitric acid, and test with chlorhydric, acid according to (64) in order to confirm the presence of SILVER.

If silver is present, repeat the treatment with potassic cyanide until no more silver is dissolved.

SULPHUR. Test the substance, free from plumbic sulphate (126) and argentic chloride, for sulphur, according to (10) If the substance is moist, it must be carefully dried by heating it in a porcelain dish over a water-bath before applying the test.

If sulphur is present, heat the substance in a covered porcelain crucible, until the sulphur is completely volatilized.

CARBON. If the substance has a *black* or *gray* color which it (137) loses, when it is heated with the blow pipe on platinum foil, carbon in some form is probably present. If carbon is present the substance, free from plumbic sulphate, argentic chloride and sulphur, should be burnt, until as much as possible of the carbon is destroyed, by heating it red-hot on platinum foil or in a porcelain crucible.

XXIV.—FUSION WITH POTASSIC AND SODIC CARBONATES AND SODIC NITRATE.

BARIC SULPHATE, SILICIC ACID, and many (138) SILICATES, and some OXIDES. Mix the finely

- (a) powdered substance, free from plumbic sulphate, argentic chloride and sulphur, and as nearly free from carbon as possible, with two parts of potassic carbonate, two parts of sodic carbonate and one part of sodic nitrate; * bring as much of the mixture as can be heated at once on the

* The sodic nitrate is added in order to destroy carbon or other reducing substances. If the substance to be analyzed appears to contain much carbon, increase the quantity of sodic nitrate. If the substance contains no carbon, the use of sodic nitrate is usually unnecessary.

platinum foil, and heat the under side of the foil with a blast-lamp until the whole mass is in a state of quiet fusion. Repeat this operation two or three times, if much substance is required for the analysis.

- (b) Detach the fused mass from the platinum foil each time, by plunging the foil, while it is hot, in distilled water. Boil the product of fusion with water, and if it does not dissolve completely, filter, and wash the precipitate on the filter with distilled water, rejecting the washings. Continue the washings, until baric chloride ceases to produce a precipitate in the water, which runs through the filter.
- (139) The first filtrate may contain ARSENIC ACID, see (116), (its occurrence is rare); CHROMIC ACID, see (117); SULPHURIC ACID, (119), (the tests, referred to above, may be applied successively to a single portion of the filtrate); FLUORHYDRIC ACID (its occurrence is rare), see (122 and (42); and PHOSPHORIC ACID, (123). The two last tests may be applied successively to another portion of the filtrate. No compound of these acids except BARIC SULPHATE is by itself insoluble, but insoluble substances sometimes contain small quantities of the acids. CALCIC FLUORIDE is only decomposed completely by the treatment with sulphuric acid described in (42).

SILICIC ACID. The principal portion of the filtrate should (140) be tested according to (60) for silicic acid. After separation of silica, the only metals,* that can be present in the chlorhydric acid solution, are LEAD, ALUMINIUM and ZINC. Test for lead by adding an excess of dilute sulphuric acid and alcohol to the solution. If a precipi-

* It is evident that sodium and potassium in insoluble silicates cannot be detected by this process. All reliable methods for their detection require the use of platinum vessels and great care in manipulation. Larger works on analysis must be consulted for such methods.

tate of PLUMBIC SULPHATE forms, filter. Test for ALUMINIUM, in the solution, free from lead, by adding ammoniac hydrate in excess. If a precipitate of ALUMINIC HYDRATE forms, filter. Test for ZINC in the solution, free from lead and aluminium, by adding to the solution containing ammoniac hydrate in excess, ammoniac sulphide. *A flocculent white* precipitate consists of SULPHIDE OF ZINC.

- (141) If a portion remains insoluble after boiling the fused mass with water (38), (b), dissolve it in chlorhydric acid. (If much silica was discovered (see (40), it is best to evaporate the chlorhydric acid solution to dryness, and to proceed as directed in (60).

Test for metals in the chlorhydric acid solution according to page 101, VI, &c.

REMARKS.

THE tables, I, II and III, contain a synopsis of the course of analysis of bodies in solution, given in Part III, and they are intended as an index to the methods which are there described in detail. They may also serve as guides in analytical work, to students who have made themselves acquainted with the detailed descriptions of Part III.

A skeleton form, similar to that of the tables, should be filled out with the results of an analysis, and the reactions which occur on the application of each test should be noted.

The sign — placed under the formula of a compound indicates that it is formed as a precipitate during a reaction. This sign is used in the following tables, and it will also be found convenient in noting the results of analyses.

The tables IV and V are intended to indicate the degree of solubility in water, and in many cases in alcohol, acids and alkalis, of the combinations of the metals and acids, mentioned in Part II.

The properties of a salt are described in the square formed by the intersection of the column devoted to an acid with that devoted to a metal.

The roman numerals, standing after the symbols of the metals, indicate their atomicity, and the formula of a salt is made by putting the symbol of a metal in the place of the symbol of an equivalent number of atoms of hydrogen in an acid. When an acid contains more than one atom of hydrogen, several classes of salts may be formed, according as one or more atoms of hydrogen are replaced by a metal. The normal or regular salts are

those which are formed by the replacement of *all* the atoms of hydrogen by a metal.

The descriptions of the tables refer to normal salts, but the following cases are exceptions, because the salts specified are more commonly met with in analysis; and in using the tables, the formulas below must be substituted for those of the normal salts:—

ARSENIATES.— $\text{MgNH}_4\text{AsO}_4$; $\text{MnNH}_4\text{AsO}_4$; $(\text{Hg}_2)\text{HAsO}_4$; HgHAsO_4 .

PHOSPHATES.— $(\text{NH}_4)_2\text{HPO}_4$; BaHPO_4 ; CaHPO_4 ; MgNH_4PO_4 ; MnNH_4PO_4 ; HgHPO_4 ; Na_2HPO_4 .

The ARSENIATE OF ALUMINIUM probably contains more acid than the normal salt.

The CHROMATES OF ALUMINIUM and of IRON (ferric chromate) contain a larger proportion of metal than the normal salts.

ARSENITES.—The arsenites referred to in the table have only *two* atoms of hydrogen replaced by a metal, except $\text{Mg}_3(\text{AsO}_3)_2$ and Ag_3AsO_3 , which are normal salts.

The ARSENITES OF COBALT and MANGANESE contain less than two atoms of hydrogen, replaced by the metal.

BORATES.— $(\text{NH}_4)_2\text{B}_4\text{O}_7$; BaB_2O_4 ; CuB_4O_7 ; FeB_4O_7 ; $(\text{Fe}_2)\text{B}_3\text{O}_6$; PbB_4O_7 ; CaB_2O_4 ; MnB_4O_7 ; NiB_4O_7 ; $\text{K}_2\text{B}_2\text{O}_4$; $\text{Ag}_2\text{B}_2\text{O}_4$; $\text{Na}_2\text{B}_4\text{O}_7$; ZnB_4O_7 .

EXPLANATION OF SIGNS IN TABLES IV AND V.

Numbers.=number of parts of water required to dissolve one part of the anhydrous salt * at the ordinary temperature.

∞.=insolubility.† The sign of infinity indicates that an infinite quantity of water is required to dissolve the salt.

s.=soluble to a considerable extent in water.

s.s.=slightly soluble.

del.=deliquescent, or capable of dissolving by attracting moisture from the air.

dec.=decomposed. *Examples:* dec.=decomposed by water. —dec.=decomposed by acids.

—=acids. *Example:* —s.=soluble in acids.

+ =sodic or potassic hydrate. *Example:* + s.=soluble in sodic or potassic hydrate

am.=ammonic hydrate.

am. cl.=ammonic chloride.

al.=alcohol. *Example:* al. ∞.=insoluble in alcohol.

When no solvent, such as —, am., al., &c., is indicated, the signs: ∞, s., s.s., and dec. refer to the action of water on the salt.

* The salt without water of crystallization is referred to.

† Most of the salts, marked insoluble in the table, are not really more insoluble than salts like baric-sulphate, but they are described as insoluble, because they are known to be nearly so, and because the quantity of water required to dissolve them has not been determined.



TABLE I.

Tests for metals

CHLORHYDRIC ACID TEST.	
<i>Metals of Group V., Section I., in acid solution : p. 98, III.</i>	
Lead. PbCl_2 , <i>white, crystalline</i> ; sol. in boiling water. The solution, $+\text{H}_2\text{SO}_4$, $+\text{alcohol}$, gives PbSO_4 , <i>white, pulverulent</i> .	SOLUBLE IN As_2S_3 ; <i>yellow</i> SnS_2 , <i>yellow</i>
Silver. AgCl , <i>white, curdy</i> ; sol. in NH_4HO . The solution, $+\text{HNO}_3$, to acid reaction, gives AgCl .	
Mercurous Salts. Hg_2Cl_2 , <i>white, pulverulent</i> ; insol. and turns black with NH_4HO .	AFTER OX. Al
<i>Metals in alkaline solution : p. 99, IV.</i>	
PbCl_2 ; PbSO_4 ; AgCl ; when the precipitate is <i>white</i> . See (66) (a)	Arsenic. Ni (81) n A —
As_2S_3 ; Sb_2S_3 ; Sb_2S_5 ; SnS ; SnS_2 ; where the precipitate is <i>colored</i> . See (66) (b)	Antimony. (82)
<i>Acids, usually in alkaline solution : p. 100, V.</i>	
(60) Silicic Acid, SiO_2 , insol. in acids after drying. See p. 95.	
(68) Carbonic Acid, CO_2 . The gas gives CaCO_3 in a drop of lime water.	Tin. SnO_2 , in (83) —
(69) Cyanhydric Acid, HCN . Smell of bitter almonds.	
(70) Sulphydric Acid, H_2S . Smell of rotten eggs. The gas blackens lead-paper.	
(71) Sulphurous Acid, SO_2 . Smell of burning sulphur. The gas turns <i>blue</i> a mixture of ferric chloride and potassic ferricyanide.	Gold, Au, in (84) — Tl <i>flu</i>

ed successively.

SULPHYDRIC ACID TEST: p. 101, VI.

Metals of Groups V. and VI., in acid solution.

C SULPHIDE: p. 102, VII.

*Sb₂S₃, orange; SnS, brown;
—
dark brown.*

**INSOLUBLE IN AMMONIC SULPHIDE:
p. 105, IX.**

HgS; PbS; Bi₂S₃ and CuS, black.

**WITH SODIC CARBONATE
E: p. 104, (79)**

Mercury. HgS, insol. in boiling HNO₃.
(86)

**SOLUBLE IN BOILING NITRIC ACID:
p. 106, (85).**

*in cold water. The solution,
in acetic acid, + AgNO₃, gives
pulverulent.*

PbS; Bi₂S₃ and CuS.

Lead. The solution in HNO₃, after evapora-
(87) tion with H₂SO₄, + water and alco-
hol, gives PbSO₄, *white, pulverulent.*

*sol. in cold water; sol. in hot
solution, + Zinc and Platinum
with each other, gives Sb,
in the platinum.*

Bismuth. The solution in HNO₃, + NH₄HO,
(88) gives Bi(HO)₃, *white, flocculent*;
insol. in an excess of NH₄HO.

*; sol. in hot HCl. The solu-
gives metallic flakes, Sn. The
precipitate, dissolved in a little
gives Hg₂Cl₂, *white, pulveru-**

Copper. The solution in HNO₃, + NH₄HO,
(89) gives a *deep blue* solution.

*and HCl; sol. in aqua-regia.
+ FeSO₄, gives Au, *purple,**

TABLE II.

Tests for meta

AMMONIC SULPHIDE TEST: p. 108, X.

*Metals of Groups III. and IV., in a solution, containing NH_4Cl , and neu*INSOLUBLE IN COLD, DILUTE CHLOR-
HYDRIC ACID: p. 109, (92)CoS and NiS, *black*.SOLUBLE IN COLD, DILUTE
p. 110,FeS, *black*; MnS, *flesh-colored*; $\text{Cr}_2(\text{HO})_3$, *dirty green*.COBALT. The insoluble sulphides give a *blue*
(93) borax-bead, if much cobalt is present. If the cobalt color is obscured by nickel, it appears after the reduction of the nickel on charcoal.The HCl solution, freed from H_2S (see 95), is boiled with NaHO

INSOLUBLE IN SODIC H

 $\text{Fe}_2(\text{HO})_6$; $\text{Mn}(\text{HO})_2$; $\text{Cr}_2(\text{HO})_3$.Iron. $\text{Fe}_2(\text{HO})_6$, dissolved in HCl
(97) a red color.NICKEL. When little or no cobalt is present
(94) the insoluble sulphides give a *brown* nickel-bead.When much cobalt is present, dissolve the sulphides in boiling dilute H_2SO_4 , + a little HNO_3 . Precipitate with potassic ferricyanide, add NH_4HO to dissolve the nickel precipitate, and add BaCl_2 to render the filtration easy. The filtrate, + $(\text{NH}_4)_2\text{S}$, gives NiS, *black*.Manganese. $\text{Mn}(\text{HO})_2$ imparts a
(98)Chromium. $\text{Cr}_2(\text{HO})_3$, after ox
(99) Na_2CO_3 , and KNO₃ tion, gives a *yellow* addition of acetic

SOLUBLE IN SODIC HY

 $\text{Zn}(\text{HO})_2$ and $\text{Al}_2(\text{HO})_3$.Zinc. The NaHO solution, + H_2S ,
(100)Aluminium. The NaHO solution
(101) then + NH_4HO to a *white, flocculent*.See p. 112, XI. for the above tests for
BASIC, AND PHOS

ed successively.

	TESTS FOR METALS OF GROUP II: p. 115, XII.
NH_4HO . (90)	Add ammoniac hydrate + ammoniac carbonate to a solution containing metals of Group II. mixed with ammoniac chloride. $BaCO_3$ and $CaCO_3$, white. MAGNESIUM compounds remain in solution.
YDRIC ACID :	
$Al_2(OH)_3$, white ;	Barium. The precipitate dissolved in dilute chlorhydric acid + $CaSO_4$, gives $BaSO_4$, white, pulverulent.
oxidized by HNO_3 , (96).	Calcium. The solution (109), free from Ba and rendered alkaline with NH_4HO , + $(NH_4)_2C_2O_4$, gives CaC_2O_4 , white, pulverulent.
110, (96)	Magnesium. The solution, freed from Ba and Ca by treatment with ammoniac hydrate and carbonate, + Na_2HPO_4 , gives $MgNH_4PO_4$, white.
alphocyanate, gives	
the soda-bead.	TESTS FOR METALS OF GROUP I: p. 117, XIII.
platinum-foil with in water and filtra-) solution, on the	Sodium. Yellow color of flame. (The higher metals must be absent.)
10, (96)	Potassium. Violet color of flame, seen through a blue glass. (The higher metals must be absent.)
white, flocculent.	Ammonium. Solution, + $NaHO$, gives NH_3 , ammonia smell.
acid reaction, and on, gives $Al_2(OH)_3$,	
ce of OXALIC, Bo-	
is.	

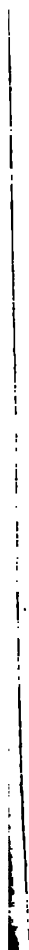


TABLE IV. EXPLANATION OF SIGNS (see p. 131); numbers = d
del. = deliquescent; *dec.* = decomposed; — = acids; + = alkalis

	Aluminium. Al. VI.	Ammonium. NH ₄ . I.	Antimony. Sb. III and V.	Barium. Ba. II.
Acetic Acid, $\text{HC}_2\text{H}_3\text{O}_2$	8,4	del. al. s.	s.	1,4 al. 100.
Arsenic Acid, H_3AsO_4	∞ . — s.	s.	∞ . HCl. s.	s.s. — s.
Arsenious Acid, H_3AsO_3	∞ . — s.	s.	dec. — s.; + s.	s.s. — s.
Boracic Acid, H_2BO_3	s.s. — s.	s.		s.s. — s.
Carbonic Acid, H_2CO_3	∞ . — s.	Sesquicar- bonate 1,7		12000. — s.
Chlorhydric Acid, HCl.....	1,4 al. s.	2,7 al. s.s.	dec. — s.	2,1
Chloric Acid, HClO_3	del. al. s.	s. al. s.		2,7
Chromic Acid, H_2CrO_4	∞ . — s.	s.		∞ . — s.
Cyanhydric Acid, HCy.....	s. al. s.	s. al. s.	∞ .	s. al. s.
Ferrocyanhydric Acid, $\text{H}_4(\text{FeCy}_6)$..		s.		s.s. — s.
Ferricyanhydric Acid, $\text{H}_5(\text{Fe}_2\text{Cy}_{12})$.		s.		s.
Fluorhydric Acid, HF.....	∞ .	s. al. s.s.	del.	s.s. — s.
Nitric Acid, HNO_3	del. al. 1.	0,5 al. 2.		12,5 al. ss.
Oxalic Acid, $\text{H}_2\text{C}_2\text{O}_4$	ss. — s.	25. al. ∞ .	∞ . — s.	2600
Phosphoric Acid, H_3PO_4	∞ . — s.	4. al. ∞ .	∞ .	20500. — s.
Sulphuric Acid, H_2SO_4	3. al. ∞ .	1,3 al. s.s.	dec. — s.	43000. — ∞ .
Sulphurous Acid, H_2SO_3	∞ . — dec.	1. al. s.s.	∞ . — dec.	s.s. — dec.
Sulphydric Acid, H_2S	dec.	s. al. s.	∞ . — ∞ ; + s.	s.
<i>Hydrate, H(HO)</i>	∞ . — s.; + s.	0,3. al. s.	s.s. — s.; + s.	20.

of solubility; ∞ . = insoluble; s. = soluble; s.s. = slightly soluble; am. = ammoniac hydrate; am. cl. = ammoniac chloride; al. = alcohol.

th. II.	Calcium. Ca. II.	Chromium. Cr. VI.	Cobalt. Co. II.	Copper. Cu. II.	Gold. Au. III.	Iron. Fe. II.	Iron. Fe. VI.
	5,6 al. 21.	s.	del.	14,8 al. 13.		s.	s. al. s.
	∞ . — s.	∞ . — s.	∞ . — s.; am. s.	∞ . — s.; am. s.		∞ . — s.	∞ . — s.
	4000. — s.		∞ . — s.; am. s.	∞ . — s. + s.		∞ . — s.; am. s.	∞ . — s.
	s.s. — s.	∞ . — s.	s.s. — s.	s.s. — s.		∞ . — s.	∞ . — s.
	70600. — s.	∞ . — s.	∞ . — s.	∞ . — s.		∞ . — s.	∞ . — s.
	1,4 al. s.	del. al. s.	s. al. s.	1,5 al. s.	1,5	2. al. s.	del. al. s.
	del.		del.	del.		s.	s.
	s.	dec. — s.	∞ . — s.	s.s. — s.			∞ . — s.
	s.	∞ . — s.	∞ . am. s.	∞ . — ∞ ; + s.	s. al. s.	∞ . — s.	
	del. al. ∞ .		∞ . — ∞ .	∞ . — ∞ ; am. s.		∞ . — ∞ .	∞ . — ∞ .
	del.		∞ . — ∞ ; am. ∞ .	∞ . — ∞ ; am. s.		∞ . — ∞ ; + dec.	s.
	27000. — s.s.	s.	s.s. — s.	s.s. — s.		s.s. — s.	s.s. — s.
	2. al. s.	2. al. s.	del. al. 1.	1. al. s.		s.	s. al. s.
	500000. — s.	s.	s.s. — s.; am. s.	∞ . — s.		4500 — s.	s.s. — s.
	∞ . — s.	∞ . — s.	∞ . — s.; am. s.	∞ . — s.		∞ . — s.	∞ . — s.
	380. al. ∞ .	s.	44.	3,6		5.	s.
	800.	∞ .	s.s. — dec.	s.s. — dec.		s.s. — dec.	s.s. — dec.
	500.		∞ . — dec.	∞ . — ∞ .	∞ . — ∞ .	∞ . — dec.	∞ . — dec.
	584 — s.	∞ . — s.; + s.	∞ . — s.; am. s.	∞ . — s.; am. s.	∞ . — s.; + s.	150000 — s.; am. s.	∞ . — s.

TABLE V. EXPLANATION OF SIGNS (see p. 131); number = *d*
del. = *deliquescent*; *dec.* = *decomposed*; — = *acids*; + = *alkalis*

	Lead. Pb. II.	Magnesium. Mg. II.	Manganese. Mn. II.	Mercury. Hg ² . II.
Acetic Acid, $\text{HC}_2\text{H}_3\text{O}_2$	2. al. s.s.	2. al. s.	5. al. s.	769.
Arsenic Acid, H_3AsO_4	∞ . — s.	∞ . — s.	∞ . — s.	∞ . — s.
Arsenious Acid, H_3AsO_3	∞ . — s.	∞ . — s.	∞ . — s.	∞ . — s.
Boracic Acid, H_2BO_3	∞ . — s.	s.s. — s.	∞ . — s.	
Carbonic Acid, H_2CO_3	50500. — s.	2500. — s.	∞ . — s.	∞ . — s.
Chlorhydric Acid, HCl	185. — s.s.	1,8 al. 7	1,5 al. s.	∞ .
Chloric Acid, HClO_3	s. al. s.	del.	s.	
Chromic Acid, H_2CrO_4	70000. — s.	s.	∞ . — s.	∞ . — s.
Cyanhydric Acid, HCy	∞ . — s.	s.	∞ . — s.	∞ .
Ferrocyanhydric Acid, $\text{H}_4(\text{FeCy}_6)$..	∞ . — ∞ .	s.	∞ .	∞ . — ∞ .
Ferricyanhydric Acid, $\text{H}_5(\text{Fe}_2\text{Cy}_{12})$.	s.s. — s.s.	s.	s. — ∞ , am. ∞	
Fluorhydric Acid, HF	s.s. — s.	∞ . — s.	∞ . — s.	∞ .
Nitric Acid, HNO_3	1,5	1. al. s.	del.	dec. — s.
Oxalic Acid, $\text{H}_2\text{C}_2\text{O}_4$	s.s. — s.	2000. — s.; am. cl.s.	3000. — s.	∞ . — s.
Phosphoric Acid, H_3PO_4	∞ . — s.	∞ . — s.	∞ . — s.	∞ . — s.
Sulphuric Acid, H_2SO_4	13000. — ∞ ; al. ∞ .	3. al. ∞ .	1,5	s.s.
Sulphurous Acid, H_2SO_3	∞ . — dec.	s.s. — dec.	s.s. — dec.	
Sulphydric Acid, H_2S	∞ . — ∞ .	s.s. — dec.	∞ . — dec.	∞ . — ∞ .
Hydrate, $\text{H}(\text{HO})$	10000. — s.; + s.	5000. — s.; am. cl.s.	∞ . — s.; am. cl.s.	∞ . — s.

m. = ammoniac hydrate; *am. cl.* = ammoniac chloride; *al.* = alcohol.

<i>r.</i>	Nickel. Ni. II.	Potassium. K. I.	Silver. Ag. I.	Sodium. Na. I.	Tin. Sn. II.	Tin. Sn. IV.	Zinc. Zn. II.
	9,6	0,4 del. al. 3.	100.	8,5 al. s.	s.	s.	4. al. s.
	∞. —s.; am. s.	del. al. s.	∞. —s.; am. s.	8,4 al. s.	∞. —s.	∞. —s.	∞. —s.
	∞. —s.; am. s.	del. al. ∞.	∞. —s.; am. s.	s.	∞. —s.	∞. —s.	∞. —s.
	∞. —s.	s.	s.s. —s.; am. s.	Borax 12.			∞. —s.
	∞. —s.	0,9 del. al. s.s.	∞. —s.; am. s.	6. al. s.s.	∞. —s.		∞. —s.
5.	4.	3. Al. s.s.	∞. —∞.; am.s.	2,6 al. s.s.	s. al. s.	s. al. s.	0,3 al. 1.
	del.	16,7 al. 120.	11.	1,			del.
	s.	2. al. ∞.	∞. —s.	s. al. ∞.	∞. —s.	∞. —s.	s.
6.	∞. —∞.; am. s.	del. al. s.s.	∞. —∞.; am.s.	s. al. s.s.	∞.		∞. —s.; +s.
	∞. —∞.; am. s.	4,5 al. s.s.	∞. —∞.; am. s.	5.			∞. —∞.; +s.
	∞. —∞.; am. s.	2,5 al. s.s.	∞. —∞.; am. s.	5,5	∞.		∞. —s.; +s.
	s.s. —s.	del. al. s.s.	del.	25.	s.	s.	s.s. —s.; am. s.
	2. al. s.	3,4 al. s.s.	0,44 al. s.	1,14 al. s.s.	dec. —s.	dec.	1. al. s.
	∞. —s.; am. s.	3. al. s.s.	s.s. —s.; am. s.	30. al. ∞.	s.s. —s.	s. —s.; am. s.	∞. —s.; am. s.
	∞. —s.	s. al. ∞.	∞. —s.; am. s.	24. al. ∞.	∞. —s.	∞. —s.	∞. —s.; am. s.
	3.	9,6 al. s.s.	200.	7,4 al. s.s.	s.	s.	2. al. ∞.
7.	∞. —dec.	del. al. s.s.	s.s. —dec.	9.			s.s. —dec.
	∞. —dec.	del. al. s.	∞. —∞.	s. al. s.s.	∞. —∞.; +s.	∞. —∞.; +s.	∞. —dec.
	∞. —s.; am. s.	0,5 al. s.	Ag ₂ O. 3000. —s.; am. s.	1,7 al. s.	∞. —s.; +s.	∞. —s.; +s.	∞. —s. +s. am.

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